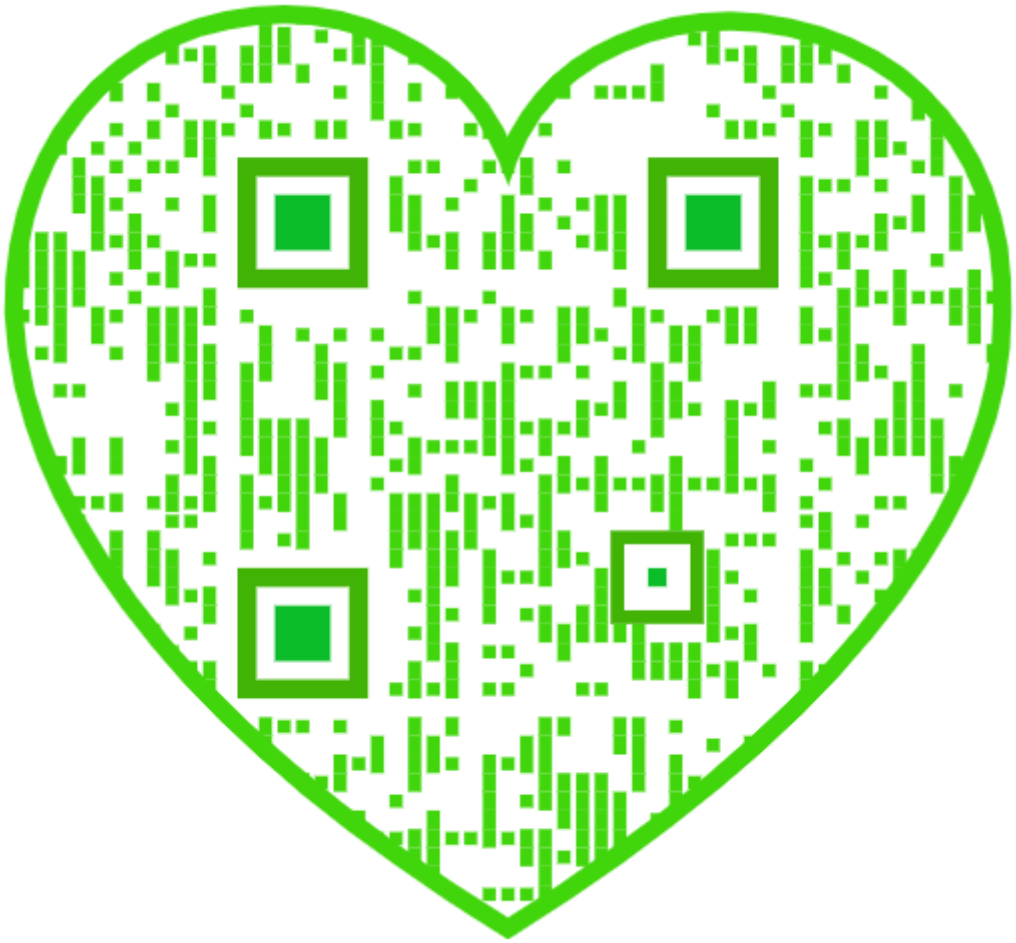

TEXT BOOK OF CHEMISTRY

FOR

NURSES AND STUDENTS OF HOME ECONOMICS

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TEXT BOOK OF CHEMISTRY

FOR

NURSES AND STUDENTS OF HOME ECONOMICS

BY

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PREFACE

The purpose of this book is to present in one volume those general principles of chemistry, inorganic, organic, and physiological, which give the necessary foundation for practical courses such as nutrition, dietetics and cookery, materia medica, bacteriology, and so forth, which are more or less directly dependent on a basis of chemical theory. With this object in view I have eliminated much that would be included in the conventional chemistry course, in order to avoid confusing the student who has only a limited amount of time to spend on this subject. On the other hand I have included certain topics which are not usually dealt with in a course so elementary in character but which lead to a fuller understanding of everyday phenomena.

Throughout the text I have kept particularly in mind the needs of the nurse-in-training, since my experience in teaching this class of student has led me to believe that they should have a rather special type of instruction in chemistry. In selecting my material I have endeavoured to conform to the requirements of the Committee on Education of the National League of Nursing Education, and of the New York State Board. Since, however, these requirements are based on a consideration of those applications which will be made in the later experience of the nurse, an experience which has many points in common with that of the Home Economist I trust that this text may prove equally useful for students of both types.

Inasmuch as many, if not most of our hospitals have neither the time nor the equipment necessary to include

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laboratory work or even lecture demonstration as part of the course in chemistry, I have endeavoured to produce a text which can be used with profit without these aids. Since, however, it is unquestionably desirable to introduce experimental work when possible, in order that the student may acquire dexterity in manipulation, accuracy in observation, and neatness and despatch in following directions, I have included a practical manual of experiments, requiring only the simplest apparatus, which may be used either whole or in part for laboratory work or class demonstration. The experiments marked Demonstration are a little more elaborate than the others, requiring more time or more apparatus than is usually convenient to provide for each member of the class, but they may advantageously be carried out by the instructor as illustrations of certain points taken up in the class work.

It is very frequently the case that some members of a class have already had a course of chemistry in high school. Such advanced students might very well be entrusted with the task of carrying out the demonstrations, preparing solutions, etc., under the supervision of the instructor, thus acquiring a little more experience and manipulative skill, instead of repeating those experiments with which they are already familiar.

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TEXTBOOK OF
CHEMISTRY FOR NURSES
AND
STUDENTS OF HOME ECONOMICS

SECTION I

INORGANIC CHEMISTRY

CHAPTER I

INTRODUCTORY

In the science of chemistry we study the behavior of substances under varying conditions and try to ascertain of what they are composed, why they behave as they do, and whether there are any general rules which govern their behavior and from which we can foretell the action of these and other substances in all circumstances. Many such rules have been discovered and are now well-established, and the study of chemistry is largely reduced to the learning of these rules and of how to apply them.

The first important generalization which we learn to make is that substances can be divided into two groups, *elements* and *compounds* If we take any solid substance we can by grinding reduce it to powder so fine that the

particles can no longer be distinguished by the naked eye. If we continue still further it is possible to get it so fine that the individual particles cannot be detected with an ordinary microscope. Eventually, by special methods, it is possible to reach a point where the separate particles can no longer be seen even with the aid of the ultramicroscope, the strongest microscope we have. Mechanical division can go no further, and yet each of these minute particles has all the properties of the original substance and they behave in a mass exactly as did the undivided whole. If we now resort to chemical means of separating our substance it will be found that some substances can be separated beyond this point into two or more different kinds of matter with distinct characteristics and different properties, and which by suitable means can be made to recombine into the original substance again. Such substances are called compounds. Water is a compound; it can be separated by means of electricity into two gases, oxygen and hydrogen, and these, if heated to a sufficiently high temperature will combine and form water again. Oxygen and hydrogen are elements. They cannot be further decomposed by any means in our power. Of all the infinite variety of substances of which our universe is made up we know of only about eighty elements; all the rest are compounds of these in different proportions.

The most important of the elements are carbon, oxygen, hydrogen, and nitrogen. Silicon indeed makes up about a quarter of the earth's surface, but as it plays only a very minor part in the life processes it is of little interest save to the geologist. The percentage of the various elements present in the human body according to a recent estimate is as follows:[1]

[1] Sherman, Chemistry of Food and Nutrition, 2nd ed. p. 234.

	Per cent.
Oxygen	65
Carbon	18
Hydrogen	10
Nitrogen	3
Calcium	2
Phosphorus	1
Potassium	0.35
Sulphur	0.25
Sodium	0.15
Chlorine	0.15
Magnesium	0.05
Iron	0.004
Iodine	
Fluorine	Very minute quantities.
Silicon	

A *mixture* of two substances differs from a compound inasmuch as there is no union between the constituents. Substances can be mixed in any proportion, can be separated again mechanically, and do not change their individual characteristics with the mixing. Salt and sand can be mixed together in equal quantities, or half as much salt as sand, or twice as much salt as sand, it makes little difference which. The product will behave like salt and sand; it develops no new properties of its own. By treating it with water, filtering, and letting the water evaporate we get our salt and our sand separate again. Powdered sulphur can be mixed with iron filings giving a mixture from which the iron and sulphur can readily be separated, and in which each substance behaves as it would if it were alone. If we heat this mixture gently however something happens; a bright red glow spreads through the whole mass, lasting even after the flame has been removed, the separate particles of iron and sulphur fuse together into a solid mass from which they cannot again be separated, and, most important of all, the product has entirely different properties from either of the constituents. The mixture has been converted by heat into a compound.

The smallest part of an element which can exist as a whole unit is called an *atom* (from the Greek *atomnos* indivisible). This conception goes back to the early days of the world's thought. The old philosophers imagined that the atoms of all elements were composed of the same kind of matter, but differed in size and shape, hence showed different properties. Those substances which are pleasant and attractive to our senses were supposed to be made up of smooth round atoms, while unpleasant substances were composed of jagged, irregular atoms which force their passage roughly into our perceptions. This view, with some modifications, held sway more or less until about the beginning of the nineteenth century, when the English chemist Dalton propounded his epoch-making atomic theory. So simple does this theory sound that it is difficult to realize that only after its adoption could the development of modern chemistry begin.

According to Dalton, the atoms of any one element are alike in all properties but differ from the atoms of any other element. A compound is formed by the combination of atoms of different kinds. The nature of the compound depends upon the number and kind of atoms present, therefore the composition of a compound must be fixed and definite, unlike that of a mixture which may vary. An atom is indivisible, cannot be broken up into anything smaller; it is the smallest part of an element which can take part in a chemical action. The smallest part of a compound, however, must consist of two or more atoms which would be separated if the compound were decomposed. The early chemists spoke of atoms of elements and compounds alike, but this led to confusion since there were then two kinds of atoms, one divisible and the other indivisible. The term *molecule* (diminutive of *mol*, a mass) was therefore adopted for the

smallest part into which a compound can be divided without losing its individual character. A molecule is the smallest combination of atoms which can exist alone. Only in a few exceptional cases can a single atom exist by itself. Usually there are two or three atoms of different kinds combined, but sometimes we find two atoms of the same kind existing together. In the latter case we have a molecule, since it is the smallest combination of those atoms which can exist alone, but a molecule of an element, since both atoms are of the same kind.

The properties of an atom are invariable and unchangeable, whether that atom is found alone or in combination. The widely different characters of the innumerable substances which make up the universe is due to the properties of the various atoms involved. Among these properties perhaps the most important is the weight of the atom. It is of course impossible to separate and weigh a single atom, but it is possible to compare the weights of the different elements and arrange them in a series according to their relative weights. For instance, an atom of hydrogen and an atom of chlorine combine together to give a molecule of the compound hydrochloric acid. If we take a given weight of chlorine we can form from it a definite amount of hydrochloric acid, which can be weighed. The total weight of the hydrochloric acid produced will be the sum of the weights of the hydrogen and the chlorine used. If we then take an equal weight of chlorine and combine it with another element, silver, each atom of chlorine will combine with one atom of silver to give a molecule of another compound, silver chloride. We would find that the silver chloride weighed more than the hydrochloric acid produced from the same amount of chlorine. If we start with 36 grams of chlorine in each case we will get 144

grams of silver chloride as compared with 37 grams of hydrochloric acid. The weight of the chlorine started with must have been the sum of the weights of all the atoms of chlorine present, therefore in equal weights there must have been the same number of atoms, and as one atom of chlorine entered into the composition of each molecule of hydrochloric acid and of each molecule of silver chloride the number of molecules produced would be the same in each case. The difference in the weights of the two products must therefore represent the differences in the weights of the atoms of hydrogen and silver, which we express by saying that the atom of silver is 108 times as heavy as the atom of hydrogen, or has an *atomic weight* of 108. By somewhat similar experiments we find the relative weights of the atoms of every element. Hydrogen is the lightest of all elements, so we describe the others as being so many times as heavy as hydrogen. The atomic weight of an element is the weight of an atom of that element compared with the weight of a hydrogen atom. The *molecular weight* of a compound is the weight of a molecule of it, and is the sum of the weights of the atoms making up the molecule.

Instead of writing the name of an element in full we frequently make use of an abbreviation. This is usually the first letter of the name of the element in question, but where two elements have names beginning with the same letter we avoid the difficulty either by using two letters, as Ba for barium and Bi for bismuth, or by taking the first one or two letters of the latin name of one of them. So I stands for iodine, but Fe (from *ferrum*) for iron; S stands for sulphur, Si for silicon, and Ag (from *argentum*) for silver. These abbreviations or *symbols* have a further significance. They stand not only for the element in general but for one atom of that

element, or for a definite weight of the element which is in exact proportion to the atomic weight. When we wish to indicate more than one atom we write the proper figure in front of the symbol, except where the atoms referred to form part of a molecule, in which case the number is written as a subscript. Thus 3O represents three atoms of oxygen, but O_2 represents a molecule of oxygen made up of two atoms, and $2O_2$ represents two such molecules. It would be incorrect to write O_4, as that would represent one molecule of oxygen made up of four atoms, which does not exist.

As molecules are made up of combinations of atoms we can express a molecule of a compound by combining together the symbols of the atoms which go to make it up. Such a combination of symbols is called a *formula*, and represents one molecule. HCl is the formula for hydrochloric acid and indicates that this compound is made up of one atom of hydrogen and one atom of chlorine. It also stands for a definite weight which is represented by the molecular weight of the compound. H_2O is the formula for water, and indicates the fact, which was learned by experiment, that water is made up of hydrogen and oxygen in the proportion of two atoms of hydrogen to one of water.

With the help of symbols and formulæ we can describe any chemical action very concisely and accurately by means of an *equation*. $HCl = H + Cl$ is a way of saying that a molecule of hydrochloric acid will break up into an atom of hydrogen and an atom of chlorine. $H + Cl = HCl$ describes the reverse process. Also, since each symbol stands for a definite weight of the element represented, we can use these equations in calculating the relative amounts of any substances involved in a chemical action. For instance, since the atomic weight of hydrogen is one and the atomic weight

of oxygen is sixteen, we can tell that if we put two parts by weight of hydrogen with sixteen parts by weight of oxygen we will get eighteen parts by weight of water. If we have a larger proportion than this of oxygen it will remain unacted upon unless we increase the amount of hydrogen as well.[1]

Besides its weight, another very important property of an atom is its combining power. One atom of chlorine will combine with one atom of hydrogen, but an atom of oxygen combines with two atoms of hydrogen to form water. Is there any rule from which we can learn how the atoms will combine together and which will help us in guessing at the formula of different compounds? Every atom has the power of holding in combination a certain number of other atoms, and this power is spoken of as the *valence* of the atom. The valence of any element is shown by the number of hydrogen atoms it will combine with or take the place of. Atoms or groups of atoms which combine with hydrogen are said to be *negative*, or to have a negative valence, while hydrogen and atoms which replace hydrogen in a compound are said to be *positive*, or to have a positive valence. A *group* is a combination of atoms which has so strong an attraction for each other that they are not readily separated but pass through ordinary reactions as a unit. In writing a formula, the atom or group with a positive valence is placed first and the negative atom or group second. So we write HCl, never ClH; NaCl, never ClNa, etc. In a whole class of compounds, the hydroxides (see page 39), we find the hydrogen apparently

[1] We must note carefully the distinction between parts by weight and parts by volume. It takes twice as much hydrogen by volume to combine completely with a given volume of oxygen to form water, but hydrogen is so much lighter than oxygen that the one volume of oxygen will be found to weigh eight times as much as the two volumes of hydrogen, and so the weight relations given above hold true.

in the position of a negative atom, NaOH, KOH, etc. This is because in these cases it forms part of the so-called *hydroxyl* group, OH, the group as a whole being negative although it contains a positive as well as a negative constituent.

In forming a compound a positive element combines with a negative element but two positive or two negative elements have no attraction for each other and are never found directly combined. The valences may be rather crudely regarded as so many arms or hands by which the atoms hold to one another. A monovalent element is one-armed, a divalent one two-armed, and a trivalent, three-armed. On paper these valences, or arms, or *bonds* as they are sometimes called, may be indicated by lines going out from the symbol of the atom, H—, O<, Fe<. A divalent atom can combine with one other divalent atom, thus Mg< + >O = Mg<>O (MgO), or with two monovalent atoms, thus,

$$2H— + >O = \begin{matrix} H \\ H \end{matrix}\!\!>O, (H_2O),$$

but it is impossible to have a divalent atom combined with only one monovalent atom, because in that case one valence of the divalent atom would be uncombined and no atom can remain in that condition. In the hydroxyl group the O has two negative valences while the H has only one positive valence, which leaves one negative valence to combine with another positive atom. In water this second valence is combined with another hydrogen, $O\begin{matrix} H \\ < \\ H \end{matrix}$, whereas in the hydroxides it is combined with some other element, such as sodium or potassium, or with the monovalent group NH_4.

For the most part, the valence of an element is un-

alterable, but there are a few elements which change their valence under different conditions, such as mercury which is sometimes monovalent and sometimes divalent, and iron which is sometimes divalent and sometimes trivalent.

The following table gives the valences of a few common elements and groups.

Monovalent	Divalent	Trivalent
H	Hg (Sometimes monovalent)	Fe (Sometimes divalent)
Na	O	PO_4
NH_4	SO_4	N (Sometimes pentavalent)
OH	CO_3	
Cl		
NO_3		

Knowing these it is easy to write the formulæ of compounds in which they occur, or, conversely, to determine the valence of any other element from inspection of the formulæ of its compounds. For instance, if we wish to know the formulæ of the oxide of sodium, since sodium is monovalent it will take two atoms of it to combine with one divalent oxygen, so the formula will be Na_2O, while the oxide of mercury may be either Hg_2O or HgO according as the mercury is mono or divalent. On the other hand, from the formula K_2SO_4 we conclude that the element potassium is monovalent since two atoms of potassium are required to combine with one divalent SO_4 group.

CHAPTER II

Oxygen

Oxygen is the most abundant of all elements. It makes up about one-half of the whole earth, and one fifth of the air. It can be obtained pure by various means, either from the air or from some of the various compounds containing it. It is a colorless gas like air in appearance but a little heavier than air. It is slightly soluble in water, more so in cold than in warm water. When a glass of water is left standing in a warm room little bubbles of gas begin to separate out and collect on the inside of the glass (see page 28). These bubbles are mostly oxygen. This property of dissolving in water is of practical importance because it is upon the presence of dissolved oxygen in the water that all aquatic life depends. Moreover, through the instrumentality of the bacteria, it brings about the purification of our rivers from the enormous amounts of sewage deposited in them. In the presence of sufficient oxygen the noxious substances are acted upon in such a way as to convert them into something quite harmless. Chemically, oxygen is very active. Even when diluted with other gases in the atmosphere we can see its effects. The burning of wood, the rusting of iron, and the drying of paint are all examples of the activity of atmospheric oxygen. Some metals are more susceptible to the action of oxygen than others; iron is very easily affected, aluminum scarcely at all at ordinary temperatures. For this reason aluminum utensils are easily kept bright, while iron rusts quickly unless covered with some kind of

protective coating, tin, zinc, or enamel, in order to preserve it.

The combination of oxygen with other substances is known as *oxidation*, and compounds of oxygen with one other element are called *oxides*. If two atoms of oxygen are present in the molecule it is a *dioxide*, if three atoms of oxygen a *trioxide*, and so on.

Oxidation may be slow, as when metals rust, or rapid. Very rapid oxidation is usually accompanied by the evolution of light and heat, and is then known as *combustion*. The name combustion is not entirely restricted to oxidation processes, being sometimes used to describe any action which goes on with evolution of light and heat, but all ordinary combustion is oxidation. We are apt to think of burning as destruction of matter, but all our conceptions of chemistry rest upon the fundamental principle that so far as our experience teaches us there can be no such thing as destruction of matter. Matter can be made to undergo many changes, it can be dissipated and so lost to us, but we know of no change in which the actual amount of matter in the universe is altered. We can neither add to it or take away from it. In burning the greater part of the combustible material is converted into gases and so disappears from view, but if we burn a candle, for instance, in such a way that the gases produced can be collected and weighed it will be found that they actually weigh more than the candle started with, and this proves true of every case of combustion. Combustion therefore must consist in adding something to the burning material; that that something is oxygen can be demonstrated in either of two ways. Sometimes we can decompose the product again into its elements as is the case with the oxide of mercury. Or we can show that if all the oxygen be removed from the air no combustion will take place.

If combustion is oxidation and there is a large supply of oxygen present in the air all the time, why does not everything which is combustible burn up of its own accord? Why, for instance, is it necessary to use matches and kindling wood to light a fire which, once started, will go on burning till there is no fuel left? For all reactions there is a certain temperature below which the action will not take place or will take place so slowly as to be imperceptible. For combustion this temperature is called the *kindling point*, and in the case of most substances is well above ordinary temperature. The velocity of any reaction increases with rise in temperature. At low temperatures combustion goes on with extreme slowness, but as the temperature is raised the rate increases and as the rate increases more and more heat is developed by the reaction itself, until this heat is sufficient to raise the burning substance to incandescence and to maintain it at a temperature above the kindling point until the whole is consumed. If we cool the substance down below this point the burning stops. If we place a wire gauze an inch or so above an unlighted gas jet from which a stream of gas is flowing and then light the gas above the gauze carefully it will burn above the gauze without becoming ignited below. This is because the wire conducts the heat away so rapidly that the gas below is not heated to its kindling point. When the gauze becomes hot however, as happens sooner or later, the lower part of the gas becomes heated, takes fire, and the whole burns. It is because of this property of wire gauze of conducting away and dissipating heat that miners can carry a lantern surrounded with such gauze into a mine full of inflammable gases. The light inside will burn safely with out igniting the gas outside, unless the gauze becomes over-heated, in which case an explosion occurs. The head of a match is made of some

substance with a very low kindling point. As it burns it generates enough heat to ignite the rest of the match and this generates still more heat. If we apply a lighted match to a large piece of wood the heat is conducted away over the surface so rapidly that the wood does not take fire. In order to make a fire, then, we surround our large sticks with thin bits of dry kindling which have a low kindling point and ignite these with a match thus generating enough heat in their burning to maintain the larger pieces at ignition temperature.

Some substances oxidise so rapidly in air at ordinary temperature that the process becomes combustion. This is what is known as spontaneous combustion. Many oils, for instance linseed oil, absorb oxygen very rapidly, giving a solid oxidation product. Linseed oil is used in painting because it absorbs oxygen and forms a tough resinous skin which holds the coloring matter in suspension and protects the material underneath from further action of the air. Paint oils, therefore, do not dry in the ordinary sense of the word, by evaporation, but by oxidation. If rags greasy with oil or heaps of greasy material be left undisturbed for some time in a badly ventilated place the heat developed by the slow oxidation may raise the temperature high enough for spontaneous combustion.

Why is oxidation accompanied by rise in temperatures? To answer this question we must consider the energy changes in chemical reactions. *Energy* is most simply defined as ability to do work. It exists in various forms, mechanical energy, heat, light, and electrical energy. These can be transformed from one into another, but can never, by any means of which we know, be either created or destroyed. When we speak, as we sometimes do, of producing energy by some means or other we mean simply that we have changed it from a

form in which it was not available for our purposes to a more useful form. For instance, we can take the mechanical energy of falling water and use it to turn a wheel which will set a mill in motion, or we can use that same energy to charge a dynamo and so get it converted into electrical energy which can be transmitted through a wire to a point many miles away and there be reconverted into mechanical energy again. All forms of energy can be converted into heat. When a piece of iron is hammered it becomes hot, due to the conversion of the mechanical energy of the blow into heat. When a current of electricity is passed through a wire the wire grows hot, due to the transformation of a part of the electrical energy into heat. Whence then comes the heat which appears when we burn a candle? We are not yet in a position to answer this question directly, but a solution may be arrived at from a consideration of certain other phenomena.

If we pass a current of electricity through water containing a little acid the water is separated into its elements, hydrogen and oxygen, which can be collected separately. The electrical energy necessary to accomplish this separation apparently disappears completely. No heat, light, or any other form of energy that we can recognize appears in its place. But if we put this hydrogen and oxygen together and heat them up to their kindling point they combine with a loud explosion to form water again. Whence comes the energy that manifests itself in the form of a mechanical shock to our ear drums which we call a noise? Since it cannot have been created in the process of combination it must have been present in the elements and set free from them when they combined. But in the process of separating these elements from their compound in the first place energy was used up. It seems a probable hypothesis therefore

that this apparently lost energy was really converted into some other form which remained concealed in the free elements until the moment of recombination. In the great majority of chemical changes heat is given off; for a few it is necessary to supply heat or some other form of energy which is used up as the change proceeds. When such heat appears in an action of any kind we believe that it has its source in a kind of energy which was bound up in the reacting substances and set free from them as a result of the action into which they enter. This energy is called *chemical energy*. When it is necessary to supply energy of any kind in order to keep an action going on, as in the case of the decomposition of water by the electric current, we assume that energy is actually being transformed into chemical energy. The new substance or substances formed must contain more energy than the substances started with, and this energy will reappear if we can reverse the change. The heat obtained from burning fuel therefore comes from the chemical energy set free when the burning substance combines with the oxygen of the air. If the action takes place very slowly this heat may be diffused away as fast as it is developed, so that it never becomes noticeable, as when iron rusts or when hydrogen and oxygen are left in contact at ordinary temperatures, but the total amount of heat produced is the same whether the action takes place rapidly or slowly.

Whence comes the energy involved in the life processes of animals and plants? Plants, with the aid of the green coloring matter, chlorophyll, which they contain are able to utilize the energy of the sun directly. With this energy they decompose the compounds of oxygen, water and carbon dioxide, which they absorb through their roots and leaves. Some of the oxygen they set free, and the rest with carbon and hydrogen goes to

build up their tissues. In this way energy is stored up, the energy of the sun converted into chemical energy which will become available if these compounds can be acted on in such a way as to set free this bound chemical energy; in other words, if we can reverse the process carried on by the plants. This is accomplished by oxidising the plant material. Animals get their energy from this stored up energy of plant substances which is set free by oxidation in the body, and this oxidation is carried on by the oxygen of the air which is breathed in. The oxygen combines with the hemoglobin or red colouring matter of the blood in loose combination which, from the ease with which it breaks up again, is almost more like solution than combination. There is also a little oxygen dissolved in the blood. This oxygen circulates through the tissues and there oxidizes the substances present in the cells. The dissolved oxygen is used up first, and as this is used more dissociates from the oxyhemoglobin, so that the blood always contains some free oxygen which the cells can use. In these oxidation processes the chemical energy set free is converted into heat or mechanical energy which serve to maintain the activity of the living body.

In order to have oxidation take place it is not always necessary to have free or gaseous oxygen present; instead it may be in the form of some compound from which it may easily be obtained. For instance, if water in the form of steam is passed over hot iron the iron takes the oxygen from the water forming iron oxide and leaving pure hydrogen. The process of taking oxygen from a compound which contains it is called *reduction*. Reduction is therefore the reverse of oxidation, and the two always go together. When one substance is oxidized another is reduced. In the experiment just referred to, the iron is oxidized at the expense of

the gas and will show its effects even when present in too small amount to be detected otherwise.

Carbon dioxide is formed when carbon or carbon monoxide burns in excess of oxygen. It is a colorless gas, very like air or oxygen in appearance, but with one striking difference in property. When a lighted splinter or a candle is inserted in a jar of carbon dioxide the flame is immediately extinguished. Carbon dioxide is a non-supporter of combustion. This, in connection with the fact that it is so heavy that it can be poured from one place to another like water, makes it useful as a fire extinguisher. Most of the chemical fire extinguishers on the market contain carbon dioxide, or else are so devised that carbon dioxide is generated from some materials when the extinguisher is in use. The heavy gas is poured on to the base of the flame and protects it from the action of the oxygen while itself offers no fuel to feed the fire.

Carbon dioxide is very soluble in water. At ordinary temperature and pressure water dissolves about its own volume of carbon dioxide, but at lower temperatures or under great pressure it can be made to dissolve a great deal more. *Soda water* is water into which carbon dioxide has been forced under a pressure of 60–150 pounds per square inch. The water is bottled up under this pressure, and when the pressure is withdrawn by the opening of the bottle the gas escapes, causing effervescence and frothing. Carbon dioxide is generated in the process of fermentation, of which more will be said hereafter, so if a fermenting liquid is bottled up and left fermenting pressure is developed and much carbon dioxide dissolves. On opening the bottle this escapes just as in the case of soda water. Aerated waters of various kinds are weak solutions of various mineral salts into which carbon dioxide has been forced in order to

make them more palatable. When carbon dioxide dissolves in water a portion of it reacts with the water to form a compound H_2CO_3, *carbonic acid.*

$$H_2O + CO_2 = H_2CO_3.$$

This is a very unstable compound and on gentle heating it breaks up into its constituents, H_2O and CO_2. On account of its relation to this acid carbon dioxide is often called carbonic acid gas.

Carbon dioxide is always present in the air to the extent of about 0.03 per cent. It is present in much larger amount in expired breath, having been produced in the body as a result of the oxidation processes going on there.

It is very easy to prove the presence of carbon dioxide in any gas. When carbon dioxide, or any gas containing carbon dioxide, is passed through clear lime water the lime water turns milky, owing to the formation in it of a very fine solid substance. Lime water is calcium hydroxide, $Ca(OH)_2$, and when carbon dioxide is brought into contact with it it reacts, forming solid calcium carbonate, $CaCO_3$, the same substance that is found in nature as limestone or marble.

$$Ca(OH)_2 + CO_2 = CaCO_3.$$

As no other gas has this effect on lime water the appearance of this milkiness is taken as proof of the presence of carbon dioxide.

Common *baking soda* has the formula $NaHCO_3$. When heated it gives off carbon dioxide and water.

$$2NaHCO_3 = Na_2CO_3 + H_2O + CO_2.$$

This may be used as a method of obtaining carbon dioxide, or the same result may be attained by treating the baking soda with an acid. In that case the reaction is a little different.

$$NaHCO_3 + HCl = NaCl + H_2O + CO_2.$$

Baking powder is a combination of baking soda with some substance which will react with it in the same way as the acid given above, producing carbon dioxide all through the dough so that it is made light and spongy. Sour milk, vinegar, etc. are sometimes used with baking soda to produce the necessary acid. It may be asked why it is necessary to use an acid at all for this purpose, since the carbon dioxide would be given off by the soda under the influence of the heat used in baking, even if there was no other substance there to react with it. The objection to this is that in that case the Na_2CO_3 which would be left behind in the food would give it a very unpleasant taste. It is necessary, therefore, to add something which will form a harmless and tasteless combination with the residue from the soda.

Calcium carbonate, which has already been referred to as limestone or marble, behaves like baking soda on heating. It gives off carbon dioxide and leaves behind calcium oxide, sometimes called *unslaked* or *quick lime*.

$$CaCO_3 = CaO + CO_2.$$

Calcium oxide, though not very soluble, dissolves to some extent in water, combining with it to form *calcium hydroxide*, $Ca(OH)_2$, the product being what is known as *lime water*.

$$CaO + H_2O = Ca(OH)_2.$$

This, as has already been pointed out, will react with carbon dioxide to re-form calcium carbonate again.

The Oxides of Hydrogen.—The most common oxide of hydrogen is water, H_2O, but under certain conditions we can get another, *hydrogen peroxide*, H_2O_2. This is somewhat like ozone, inasmuch as it readily gives up its second atom of oxygen. It is therefore an excellent

oxidizing agent. Pure hydrogen peroxide is a liquid, very much like water in appearance, but so strong that it blisters the skin if dropped on it. The peroxide of hydrogen used for medicinal purposes is a solution of 3 per cent. hydrogen peroxide in water. On account of its oxidizing action it decolorizes a great many colored substances and is much used as a bleaching agent for delicate materials that would be injured by strong reagents, and for the same reason it finds extensive use as a disinfectant. The advantage about its use in both cases is that the only residue left over from the reacting hydrogen peroxide is a perfectly harmless one, water.

$$H_2O_2 = H_2O + O.$$

The monoxide of hydrogen, water, is so important a subject as to need a chapter to itself.

CHAPTER III

WATER

Water plays a part in a vast number of chemical reactions, either as one of the reacting substances or as a by-product of the reaction or, sometimes, simply as the medium in which the action is carried on. A reaction between water and some other substance, in which the molecule of water separates into H and OH and the other substance also separates into two parts one of which combines with the hydrogen and the other with the hydroxyl to form two new compounds, is called *hydrolysis* (from the Greek *hydor*, water, and *lysis*, loosening). Many examples of this type of reaction will be met with later, especially in connection with the decomposition of the food substances in digestion.

Perhaps the most valuable property of water is its power of dissolving a great many other substances and so bringing them into a condition in which they are very susceptible to reaction. Every one is familiar with the phenomenon of *solution*. Salt, sugar, and other things dissolve in water, *i.e.*, the solid disappears and a clear liquid is left which may look like pure water but which has taken on new properties, for instance, in the cases mentioned has acquired a salt or sweet taste. What actually takes place when a substance dissolves in water is somewhat of a puzzle to chemists. There are those who think that there is actual combination of some kind between the dissolved substance (*solute*) and the solvent; others think that molecules of the solvent merely *diffuse* among the molecules of the solvent somewhat as

a group of people might drift into a crowd and there become separated from one another by the pressure of the people around them. Certainly if there is combination at all it is not ordinary chemical combination resulting in the formation of a chemical compound, since the proportions of a solution can be altered without altering the properties. It is easy to show that diffusion does actually take place, whether accompanied by combination or not. If a colored substance is placed in the bottom of a glass and water is added carefully and the whole left standing undisturbed for a time the colored substance will travel more or less rapidly up through the solution in opposition to the force of gravity. This diffusion is the result of a tendency on the part of the molecules to equalize the concentration or strength of the solution throughout by travelling from the part of the solution where the concentration is high to the part where it is low. The same thing is seen if a lump of sugar is dropped into a cup of tea or glass of lemonade and allowed to stand without being stirred. The sugar dissolves and forms a very concentrated sweet layer at the bottom which slowly makes its way up through the solution. If it is left long enough it will eventually become uniformly sweet all through, but this takes a long time unless the process of diffusion be hastened by stirring.

The force with which the molecules of a dissolved substance diffuse through the solution is called *osmotic pressure.* If we put in the path of the moving particles a partition through which they cannot pass but through which the particles of the solvent can pass readily, the particles of solute will exert a pressure on the walls of the partition which can be measured. If the solutions on either side of such a partition, which is called *semi-permeable,* are of unequal concentration the two will try to equalize themselves. Since this cannot be done

by the passage of the excess of solute from one side to the other it will be brought about in the only way possible, by the passage of water from the more dilute to the more concentrated, and the force with which the water flows in through the partition will be exactly equal, though opposite in direction, to the force with which molecules of solute attempt to pass out. Two solutions which are exactly equal in concentration, so that there is no tendency for water or solute to pass from one to the other, are said to be *isotonic*. If one is more concentrated than the other, the more concentrated is said to be *hypertonic*, the less, *hypotonic*. The cell wall of plant and animal cells is semi-permeable, so that while water and certain substances can pass in or out quite readily the cell contents cannot pass out. In this way a certain relation is maintained between the cell sap and the fluid which bathes the cell without. If the external fluid be greatly diluted water will enter the cell and dilute up the cell sap, and if this is allowed to go on without check the cell will swell up and eventually burst. On the other hand, if the cell be placed in a solution which is more concentrated than the cell contents water will pass out from the cell, and the contents will shrink up and draw away from the wall. This shrinkage, due to the withdrawal of water, is called *plasmolysis*, and the cell may or may not recover from it, according to the extent to which it is carried.

The corpuscles of the blood are small cells floating in the plasma or fluid part of the blood, which is a watery solution of salts along with certain other constituents. The corpuscles also contain a certain amount of water and salts and the osmotic pressure inside the corpuscle must be equal to that of the plasma outside so that while the water molecules can pass either into or out of the cell the exchange is equal in the two directions and

the condition of the corpuscle remains unchanged. If the corpuscle is placed in a more concentrated solution it will shrink, owing to the flow of water from the inside out, and if placed in a solution which is considerably more dilute and therefore has a much lower osmotic pressure than the cell contents it swells and finally ruptures owing to the increased amount of water which it has taken up. It is necessary therefore in injecting liquids into the circulation or in diluting blood outside the body to be careful to choose solutions which are either isotonic or hypotonic to the blood plasma. It has been found that a solution of from 0.7 to 0.9 per cent of sodium chloride has no effect on the corpuscles and such a solution is frequently used under the name of "*normal saline*," or "*physiological saline.*" While this is satisfactory for most purposes it does not bear any resemblance to blood plasma except in regard to the osmotic pressure since the plasma contains compounds of calcium, magnesium, potassium, etc., in addition to sodium. For this reason a preparation known as "*Ringer's solution*" is preferable to normal saline for certain purposes since it contains small amounts of calcium and potassium chlorides in addition to sodium chloride.

When a substance will not dissolve in a solvent we say that it is *insoluble* in that particular solvent, but between very soluble and absolutely insoluble substances there is a long range. The solubility of a substance in a given solvent depends partly on the nature of the solvent, and partly on the nature of the substance, and partly on the temperature employed. Some substances are almost insoluble in cold water, but readily soluble in hot. The majority of substances are more soluble at higher temperatures, but some, like sodium chloride, are very little affected by change in temperature, and a few, like calcium hydroxide, are more soluble at low than at high

temperatures. Gases belong in this latter class; hence the little bubbles formed on the inside of a glass of water which has stood in a warm room. The gases dissolved in the water at a lower temperature separate out as the temperature rises. For every temperature there is a maximum amount of any substance that will dissolve in any given solvent at that temperature.

When it is desired to get a substance into solution it is best to powder it as finely as possible in order that there may be as much surface as possible for the solvent to act upon; to stir in order to keep fresh portions of the solvent in contact with the dissolving substance continually; and, if necessary, heat. A solution which contains as much as it can hold of a solute is said to be *saturated* with that solute. If more solid is added to a saturated solution nothing happens. If you saturate a solution with a solute at a high temperature and then cool it down to a lower temperature, some of the solute will separate out, because the solvent cannot hold in solution as much solid at the lower temperature as it could at the higher. Or, if a solution is saturated and some of the solvent is then removed by evaporation solid will separate out, usually, if the evaporation has been carefully carried out, in crystalline form. In a *crystalline* solid the particles are arranged in definite geometrical form bounded by plane surfaces. When heated, crystalline substances melt at a definite temperature, unlike *amorphous* (non-crystalline) substances, which melt gradually while the temperature rises.

Many substances in the process of crystallization combine with one or more molecules of water. These compounds are comparatively unstable, losing this water readily, and with it losing their crystalline form, although their properties otherwise remain unchanged. For this reason such combined water is called *water of crystalliza-*

tion. Ordinary gypsum, calcium sulphate, crystallizes with two molecules of water of crystallization, the formula being written thus, $CaSO_4.2H_2O$, to show that the water is not an essential part of the compound. On heating it loses nearly all of this water, forming the white powder known as "Plaster-of-Paris," with the formula $(CaSO_4)_2.H_2O$. When this is mixed with water it readily takes up water again, crystallizing in minute crystals which are interlaced into a firm rigid mass which expands slightly as crystallization proceeds. If the wet mixture is poured into a mold and allowed to "set" it will be found to have completely filled all the crevices of the mold, giving a very sharp clear outline when the mold is removed. Plaster-of-Paris finds extended use in the making of surgical casts. When using it for this purpose it is important to remember that expansion accompanies crystallization, as a cast which when wet fits quite comfortably may expand and therefore tighten during setting to such an extent as to seriously impede circulation.

If we shake up an insoluble substance like sand with water it quickly settles down to the bottom again. The finer we grind the solid the longer it takes to settle, and if we have a very finely divided solid it may remain suspended in the liquid for hours or even days before it settles. Such a mixture of fine solid particles suspended in a liquid is called a *suspension.*

A suspension of fine droplets of one liquid in another in which it is insoluble is called an *emulsion.* If a liquid fat like olive oil is shaken very hard with water it forms an emulsion for a few moments, but as soon as the shaking is stopped the fat begins to separate from the water and very soon the two layers are perfectly distinct again. If a little sodium carbonate, soap, or certain other substances are dissolved in the water, however,

the fat becomes so finely divided on shaking that the separate drops are indistinguishable and the whole liquid takes on a milky appearance. An emulsion made in this way will remain for weeks, months, or even years without separating into layers, and is called a *permanent* emulsion as distinguished from a *temporary* emulsion which quickly separates into its components. Milk may be called a permanent emulsion although the greater part of the fat rises slowly to the top when the milk is left undisturbed. An emulsion of cod-liver oil is sometimes used medicinally, and the emulsification of the fats of the food by the bile is of importance in digestion.

The process of *distillation* is frequently resorted to as a method of separating pure water from other substances which may be dissolved or suspended in it. If water is heated to 100° it is converted into steam, and if this steam is passed over a cold surface it condenses to water again. The flask, or "still," in which the water is boiled is connected with a condenser, or tube surrounded by a larger outer tube through which a current of cold water circulates. By this means the inner tube is kept cold enough to condense the steam and the resulting "distilled water" runs down through the tube and out at the end. Any solid material which was dissolved or suspended in the water will remain behind in the still. It is possible in the same way to separate more or less completely two liquids whose boiling points do not lie too close together, and the process of distillation is a common method of purifying other liquids besides water. The distilling of whiskey consists in boiling off and condensing a mixture of alcohol and water, along with a small quantity of volatile flavoring material. In *fractional distillation* a mixture of two or more liquids is distilled; as the more volatile liquid boils off the tem-

perature of the vapor gradually rises, and by collecting the product (the "distillate") in separate portions or fractions, according to the temperatures at which they come off, the constituents are separated. An excellent example of this is seen in the distillation of petroleum, where crude oil from the oil wells is separated into many different fractions of quite distinct character, such as gasoline, kerosene (or coal oil), lubricating oils, and so on. *Destructive distillation* consists in heating an organic substance to its decomposition point out of contact with air so that oxidation may not take place. The gaseous and liquid products of decomposition distil off, and the solids remain in the still. *Evaporation* differs from distillation in that in evaporation the volatile vapors are allowed to escape and the non-volatile residue is preserved.

A suspension of insoluble material can be separated into solid and liquid by *filtration*, that is by pouring it carefully through filter-paper in a funnel. The liquid will pass through the pores of the paper, while the solid is retained by it. If the suspended material is very fine it will sometimes run through the filter with the liquid, but this difficulty can usually be overcome by heating the suspension for a time. The heat causes the small particles to clump together into larger lumps which can then be filtered out.

Intermediate between suspensions on the one hand and true solutions on the other are *colloidal solutions*. A colloidal solution resembles a true solution in appearance and in the ease with which it passes through filter paper, but it can be demonstrated that it is in reality a suspension of particles so fine as to approach in some cases molecular dimensions. Although these particles are quite invisible to the eye, even with the aid of an ordinary microscope, they can be detected by means of

the ultramicroscope. Substances which give true solutions with water, crystallize from water in definite crystalline form, and when dissolved in water pass through animal or vegetable membranes, are known as *crystalloids;* while substances which give colloidal solutions with water, separate in amorphous state and are unable to pass through animal or vegetable membranes, are called *colloids.* It is a well known fact, however, that substances which are normally crystalloids can at times act like colloids, while some colloids have been obtained in crystalline form. It is therefore more correct to speak of substances in the colloidal state rather than of colloids as a definite class, although the latter term is in general use and may be permissible if used with full understanding of its significance.

While some colloidal substances, such as very finely divided metals and certain metallic compounds, resemble very fine suspensions, others, such as glue, gelatine, etc., are more closely related to emulsions; they are therefore classed as *suspensoids* and *emulsoids* respectively. These two classes of colloids differ in several respects. Suspensoids show a strong tendency toward coagulation or clumping together of their particles, so that they may quite easily be induced to precipitate in an insoluble mass; emulsoids on the other hand have very little of such tendency. They can be separated from solution (as happens in the setting of glue, jelly, etc.) but readily redissolve on the addition of water. Suspensoids are very sensitive to the influence of small quantities of substances of the nature of common salt, which cause them to precipitate from solution; emulsoids are not only not precipitated themselves, but by their presence are able to protect suspensoids from precipitation. It is therefore possible by the addition of gelatine, for example, to render a sus-

pensoid so stable that it can be evaporated to dryness and redissolved again like an emulsoid. Advantage is taken of this fact in the preparation of *collargol*, a powerful bactericide. A colloidal solution of silver (suspensoid) has sufficient emulsoid added to prevent its coagulation. The water is then evaporated away, leaving the metal as a fine powder which on the addition of water passes into solution again. The effect of gelatine in ice-cream depends on the same principle. Comparatively small quantities impart to the ice-cream a smooth velvety texture instead of the usual sandy grain. This is due to the protective action of the emulsoid gelatine preventing coagulation of the casein of the milk into large lumps. For the same reason small quantities of gelatine, gruel, and so on are sometimes added to the milk for babies or invalids with weak digestion. By this means the formation of a coarse curd in the stomach is prevented, and it has been found by experiment that milk with this addition is better assimilated than it would otherwise be. In somewhat the same way gum arabic when added to candy prevents the crystallization of the sugar and gives a product with a soft, smooth texture.

A suspension of finely divided clay behaves like a suspensoid, inasmuch as the rapidity with which it settles down depends on the presence or absence of salts and emulsoid material in the water in which it is suspended. The organic matter of the soil, known as "humus," is emulsoid and therefore the water of rivers draining through land rich in humus is always muddy. Sewage is rich in colloidal material and a clear sparkling river soon becomes muddy if sewage is discharged into it in any quantity. When such muddy rivers reach the sea the suspended clay is precipitated by the salt of the sea water, resulting in the gradual filling up of the river mouth.

There is always a certain amount of attraction between two surfaces which are brought into very close contact, especially between a liquid and a solid. This is illustrated by the slowness with which the last drops of a liquid will drain out of a dish. The greater the amount of surface exposed the stronger the attraction. Every housekeeper knows that it is wasteful to put a small amount of cream, for instance, in a large pitcher because so much of it will stick to the surface of the pitcher. If the pitcher were half filled with small stones still less of the cream could be drained out, and the smaller the stones the more cream would stick to them. Colloids, which have an enormous amount of surface exposed in proportion to the amount of material, and porous substances like charcoal, have a great attraction for certain kinds of dissolved and suspended substances, and this property finds many practical applications. The use of charcoal in filters for purifying water depends on the power of the charcoal of absorbing and retaining all the impurities of the water. The use of colloidal ferric hydroxide, the so-called "dyalized iron," as an antidote in arsenic poisoning depends on the fact that the iron collects the arsenic on its surface and carries it out of the system. More important still, it is to the colloids in the soil that the latter owes its astonishing power of purifying the water which drains through it. The most remarkable illustration of this power is seen in the sewage farms, where the sewage is pumped out onto the land and allowed to filter through the soil which retains and is enriched by the organic matter while the water which drains away is so pure that it could be drunk with safety.

CHAPTER IV

Acids, Bases, and Salts

We are all familiar with the general property of *acidity*. This property is peculiar to a large group of chemical substances known as acids, or to substances containing these acids, and is generally associated with a sour taste. When we say a fruit is very acid we mean that it is very sour, and we call candies strongly flavored with lemon or lime-juice "acid drops." There are however other properties besides this sourness which are equally characteristic of acids. One, which makes it possible to detect their presence without resorting to taste, is the power of changing the color of a vegetable dye called *litmus* from its neutral shade of mauve to a bright red. They all contain hydrogen, and when their solution in water is treated with a metal such as iron or zinc hydrogen gas is produced. On evaporating the solution after the evolution of gas has ceased a new compound called a salt[1] is obtained, in which the metal has taken the place of the hydrogen set free. So invariably is this true that we derive the following definitions from this reaction.

An *acid* is a substance containing one or more atoms of hydrogen which can be replaced by a metal.

A *metal* is a substance which can replace hydrogen in an acid.

A *salt* is the compound formed when the hydrogen of an acid is replaced by a metal.

[1] "Salt" as popularly used refers to sodium chloride, NaCl, also called table salt or common salt, since it is by far the best known and most widely used member of this class of compounds.

The names and formulæ of a few typical examples of each class follow:

Acids	Metals	Salts
HCl, Hydrochloric acid.	Na, Sodium.	NaCl, Sodium chloride.
HNO_3, Nitric acid	K, potassium.	$CaSO_4$, Calcium sulphate.
H_2SO_4, Sulphuric acid.	Ca, Calcium.	Na_2SO_4, Sodium sulphate.
H_2CO_3, Carbonic acid.	Mg, Magnesium.	$Fe(NO_3)_2$, Ferrous nitrate.
H_3PO_4, Phosphoric acid.	Fe, Iron.	$FeCl_3$, Ferric chloride.
		$Ca_3(PO_4)_2$, Calcium phosphate.

Some acids contain a large amount of hydrogen, not all of which can be replaced by a metal *e.g.* $C_6H_8O_7$, citric acid, the acid present in orange and lemon juice, of which only three hydrogens can be replaced by metals, the other five remaining unchanged. Other compounds rich in hydrogen, such as turpentine, $C_{10}H_{16}$, have no hydrogen that can be so replaced, and such compounds have neither a sour taste nor the power of turning litmus red. There must therefore be some fundamental distinction between acid hydrogen and other hydrogen. Before we can attempt to determine the nature of this difference however, we must become familiar with one or two other types of compounds, of which we will first consider the salts.

As may be seen from the formulæ given above, salts are composed of a metal and all of an acid except the acid hydrogen. The residue of the acid, apart from the acid hydrogen, is alled the *acid radicle.* It cannot exist alone in the free state, but it passes unchanged from one compound to another in the course of reaction. An alternate definition to that given above for a salt is therefore: A compound of a metal with an acid radicle. The formulæ of both acids and salts are always written so as to indicate clearly what acid radicle they contain, *e.g.*, HNO_3, $Ca(NO_3)_2$, (rather than CaN_2O_6), H_2SO_4, $Fe_2(SO_4)_3$ (rather than $Fe_2S_3O_{12}$), and usually $H_3C_6H_5O_7$ in preference to $C_6H_8O_7$.

In naming the acids and salts it is desirable that the names should indicate the composition of the compound. The names of the commonest acids are made by adding the termination *—ic* to the name of the most characteristic element which it contains. H_2SO_4, sulphuric acid. H_2CO_3, carbonic acid. HNO_3, nitric acid (instead of the more awkward word nitrogenic). H_3PO_4, phosphoric acid (instead of phosphorusic). $HClO_3$, chloric acid (instead of chlorinic). Where we have two or more acids differing only in the amount of oxygen in the acid radicle the following rules for naming have been adopted. An acid exactly corresponding to an *—ic* acid, but containing an atom of oxygen less in the molecule has its name ending in *—ous*, H_2SO_3, sulphurous acid; $HClO_2$, chlorous acid, and so on. If it contains still less oxygen than an *—ous* acid the prefix *hypo-* is attached to the *-ous* name. HClO is hypochlorous acid. If it contains no oxygen at all the prefix *hydro-* is attached to the *ic* name, *e.g.*, HCl, hydrochloric acid, HBr, hydrobromic acid. On the other hand, if it contains more oxygen than the *-ic* acid this is indicated by prefixing *per-* to the name of the *-ic* acid, $HClO_4$ is perchloric acid.

Salts are named from the acids from which they are formed. Salts of *-ic* acids have the ending *-ate* substituted for *-ic*; for instance, the salts of nitric acid are called nitrates. Salts of *-ous* acids have the ending *-ite* substituted for *-ous*; the salts of nitrous acid being accordingly called nitrites. Salts of hydro-acids have names ending in *-ide*. In a few cases, of which the salts of sulphuric and phosphoric acids are the most important, the characteristic endings are attached to the first syllable of the name of the acid. So we speak of sulphates and sulphites, phosphates and phosphites, instead of sulphurates, phosphorates, and so on. When we wish to indicate the salt of some particular metal the name of

the metal is placed first, the name representing the acid radicle second; NaCl, sodium chloride, K_2SO_4, potassium sulphate. In the case of metals like mercury and iron which can exist with higher or lower valences this is indicated by the terminations *-ous* and *ic* attached to the name of the metal; HgCl, in which the mercury is monovalent, is mercurous chloride, while $HgCl_2$, in which the mercury is divalent, is mercuric chloride. Iron may be either di- or tri- valent, so $FeCl_2$ is ferrous chloride, and $FeCl_3$ is ferric chloride, and similarly for the whole series of salts of both metals.

In the formation of salts sometimes only a part of the acid hydrogen is replaced by the metal. Such salts, which still contain one or more atoms of acid hydrogen are called *acid* salts, or have the prefix *bi-* attached to the name of the acid radicle present. So $NaHCO_3$, in which only one of the two acid hydrogens of the carbonic acid is replaced by sodium is called sodium bicarbonate to distinguish it from Na_2CO_3, sodium carbonate, in which both hydrogens are replaced. Acids like phosphoric acid, in which there are three replaceable hydrogens can give three series of salts according as one, two, or all three hydrogens have been replaced. These are known as *primary*, *secondary*, and *tertiary* phosphates respectively, or may be named by putting the prefixes *mono*, *di*, or *tri*, before the name of the metal. So NaH_2PO_4 may be called monosodium phosphate or primary sodium phosphate, while Na_2HPO_4 is disodium phosphate or secondary sodium phosphate. While these rules of nomenclature may seem somewhat complicated at first sight, a little practice in their application will be sufficient to fix them in the memory.

The bases make up a third large and important class of compounds. They all react with litmus, but turn it blue instead of red as acids do; their solutions have a

more or less soapy feeling; and they all contain the *hydroxyl group*, OH, combined with a metal. The most common bases are sodium hydroxide, NaOH, potassium hydroxide, KOH, ammonium hydroxide, NH_4OH, and calcium hydroxide, $Ca(OH)_2$. They may be regarded as derived from water by replacing one hydrogen in the molecule by a metal. Their most important property is that of reacting with acids to form a salt and water, the hydrogen of the acid combining with the hydroxyl of the base to form water, while the metal and the acid radicle go together to form a salt.

$$NaOH + HCl = H_2O + NaCl.$$

$$Ca(OH)_2 + H_2CO_3 = 2H_2O + CaCO_3.$$

This reaction between acids and bases is known as *neutralization.*

The terms *base* and *alkali* are sometimes used interchangeably although they are not quite synonymous. A base is a substance which answers in every particular to the description given above, while an alkali may or may not contain hydroxyl but on dissolving in water gives a solution with basic properties. All soluble hydroxides are bases, and all bases may be called alkalies, or be said to be alkaline, but some substances which are not hydroxides, such as sodium carbonate, Na_2CO_3, are also called alkalies because their solutions in water show the properties of a base. This is due to the fact that they tend to react with the water to some extent, producing a basic solution. Sodium and potassium hydroxides which are strong bases with a caustic or corrosive action on the skin are sometimes called *caustic alkalies*, while sodium and potassium carbonate are known as *mild* alkalies as their action is much less strong. These last are also occasionally spoken of as *fixed* alkalies, to distinguish them from ammonium carbonate which is volatile.

Every atom of acid hydrogen will combine with one hydroxyl from a base. Accordingly acids with only one replaceable hydrogen are called *monobasic* since they will react with only one basic OH; those with two replaceable hydrogens are called *dibasic*, and those with three are *tribasic*. Similarly, we may speak of *mono-*, *di-*, and *tri-acidic* bases, according as they react with one, two, or three atoms of acid hydrogen.

CHAPTER V

Electrolytes and Ionization

Electricity is a powerful force of which we know little except its effects and the fact that it can be conveyed from one point to another, passing through certain substances which are called *conductors* much as water passes through a pipe. Metals such as copper, iron, and so on, are among the best conductors, that is, they offer very little resistance to the passage of the current. Certain other substances, such as glass or rubber, offer so much resistance that the current will not flow through them at all. They are therefore called *non-conductors* or *insulators*. The human body is a conductor, so if a man touches one of the poles of an electric battery he gets a shock, more or less severe according to the strength of the battery; that is, according to the amount of electricity which it produces. If his hand is protected with a thick rubber glove, or, better, if he handles the wires with a glass rod, he will feel no effect as the current is unable to pass through the insulating material into his fingers.

Electricity appears to be of two kinds, which, for convenience sake, are distinguished as *positive* and *negative*, often represented by the signs + and −. Any body which contains electricity, or from which an electric current will flow, is said to be charged with electricity. A charge may be produced by various means; as a result of chemical action, when the chemical energy of the reacting substances is transformed into electrical energy, as is seen in the storage battery in common use,

or by friction, when mechanical energy is converted into electrical energy. A simple example of this last may be seen when the hair is combed on a very dry cold day. The separate hairs stand out from the head and crackle, and sometimes even send out visible sparks. Bodies are said to be positively or negatively charged according as the electricity which they contain is positive or negative, but whenever a charge of one kind is developed on one body an equal amount of the opposite kind is developed upon some other body at the same time. It is a curious fact that two bodies charged with positive electricity or two bodies charged with negative electicity will tend to repel each other, whereas bodies charged with opposite kinds of electricity have a strong attraction for each other. In the example given above the hairs are charged with one kind of electricity, while an equal amount of the opposite kind is developed upon the comb, and accordingly each separate hair stands out as far from its neighboring hairs as possible but tries to cling to the comb.

If an electric battery is used as a source of electricity the two wires, or *poles* as they are called, leading from the battery will be charged, one with positive and the other with negative electricity. If they are connected by some conducting material a current will flow through the circuit thus made, but if a non-conductor is placed between the two poles no current passes. This can be shown by introducing into the circuit a measuring instrument known as a galvanometer which indicates by means of a pointer moving over a scale the force of the current passing between two points.

If the poles of the battery are dipped into pure water no current will pass between them, as pure water is a non-conductor, and similarly if perfectly dry salt be placed between the poles the current is interrupted.

But if a little salt is dissolved in the water the galvanometer pointer will move, showing that a current is now passing. Substances which when dissolved in water are able to carry a current of electricity are called *electrolytes*, and it has been found that acids, bases, and salts alone have this power.

The passage of electricity through a solution of an electrolyte might at first be supposed to be exactly like its passage through a metallic conductor such as a telephone wire. There is, however, one important difference. The metal remains unchanged by the passage of the current, whereas in the case of an electrolyte conductance is always accompanied by chemical action, the electrolyte being decomposed. In order to explain this, and at the same time to answer the question, How is the current carried? a theory known as the *ionic theory* has been devised which has thrown light upon a great number of otherwise inexplicable phenomena.

When acids, bases, and salts are dissolved in water they are believed to be separated by the water into particles called *ions*, each of which carries an electric charge. An ion may be a single atom or a group of atoms, and owes its peculiar properties to the charge which it carries. Acids separate into positive hydrogen ions, represented by putting a + sign above the symbol for hydrogen, H^+, and acid radicles with a negative charge, represented by a − sign, Cl^-, NO_3^-, and so on. Bases give positive metal ions and negative hydroxyl ions, *e.g.* NaOH gives Na^+ and OH^-. Salts being formed by the combination of positive metals and negative acid radicles separate into these ions again, so NaCl dissociates into Na^+ and Cl^-. An ion which is negatively charged will be attracted toward the positive pole of the battery, the positive electrode or *anode* as it is called, and will move through the water towards it, while a

positively charged ion on the other hand will travel away from the anode towards the negative electrode or *cathode*. Positive ions are therefore called *cations* and negative ions are called *anions* to indicate the direction in which they move through the solution. The amount of electricity carried by an ion is in direct proportion to its valence. A monovalent atom or group carries one unit charge, a divalent atom two such charges, and trivalent atoms three.

If a positively and a negatively charged ion happen to come into sufficiently close contact the charges neutralize each other and a molecule is formed. On the other hand there is always a tendency on the part of the water to separate the molecules into ions again, so that in any solution there is a continual breaking up of molecules into ions and recombining of ions into molecules. These two actions go on at precisely the same rate so that the sum total of the ions and molecules present remains always the same unless their equilibrium is disturbed by the addition or removal of water or salt. When an anion reaches the anode or a cation the cathode it gives up its charge and becomes an atom (or group) again, another ion being always discharged simultaneously at the opposite pole. This would tend to reduce the number of ions in the solution were it not that the balance is immediately restored by the breaking up of another molecule into ions again to replace those discharged at the electrodes.

Under ordinary conditions the molecules of a dissolved substance are never all dissociated into ions, the proportion between ionized and unionized molecules depending partly on the nature of the substance in solution and partly on the strength of the solution. As one might expect, since it is the water that causes this dissociation to take place, the more dilute the solution the greater

the number of molecules broken down into ions. On the other hand, in solutions of equal strength some substances are highly ionized while others are only very slightly ionized. All salts are highly ionized, as are also the three common acids, sulphuric, hydrochloric and nitric, and the bases sodium and potassium hydroxides. All other acids and bases are much less ionized. Since it is precisely those acids and bases which are much ionized that are known to be "strong," that is to display most markedly the acid or basic characteristics, while those that are little ionized are weak, and moreover since neither acids nor bases if perfectly free from water will conduct a current, act on litmus paper, or show any of the properties which we expect of them, we are led to believe that these properties are entirely due to the ions and not to the molecules. That is, the so-called acid properties are properties of the hydrogen ion; when there are many of these ions present in the solution it will be strongly acid; the fewer there are the weaker does the acid appear. Similarly, the basic properties of hydroxides are due to the hydroxyl ion. Since all acids give hydrogen ions and all bases give hydroxyl ions we would expect a great similarity between the different members of these two groups, which we find to be the case. Neutralization, in the light of the ionic theory, is to be looked upon not as a reaction between a molecule of acid and a molecule of base so much as a reaction between a hydrogen ion and an hydroxyl ion to form water; the metal from the base and the acid radicle from the acid remain in the solution for the most part as separate ions until evaporation of the solution brings them so close together that they unite to form molecules of salt.

In all reactions in which salts are concerned it is the ions rather than the molecules of the salt which take part.

An ion, like an atom, always maintains its individual characteristics from whatever source it comes, so the SO_4 ion, for example, will always act in the same way whether it comes from the dissociation of a molecule of H_2SO_4 or a molecule of Na_2SO_4. This simplifies the study of chemistry considerably, as we can learn certain reactions of sulphates in general or of sodium salts in general instead of learning the properties of each member of the group.

Electrolytes play an important part in the life processes. They are the most effective agents in maintaining the osmotic pressure of the body fluids and so regulating the flow of substances into and out of the cells. That this is not their sole function however is indicated by the fact that solutions of equal strengths of different salts are not exactly equivalent physiologically. The various ions have probably specific functions as well as a combined effect. A few of these special functions are recognized already, although much remains to be learned about them. It has been found for instance that the normal beating of the heart is dependent partly on the calcium ions present and partly on the sodium and potassium ions. Sodium salts are known to exercise a pronounced influence on the irritability of the muscles. Calcium ions are essential for the clotting of the blood and are sometimes added with beneficial effect in certain cases where, owing to the abnormal composition of the blood, clotting will not take place naturally and there is in consequence excessive loss of blood from the slightest wound or abrasion of the skin.

An important function of vegetables and fruits in the diet is to supply the necessary salts and acids for the body, and when these are too restricted in amount ill effects very soon make themselves felt. Sodium in the form of sodium chloride is added to the food as seasoning

and is therefore always present in sufficient, if not excessive amounts. Potassium salts are present in meat and in many vegetables and these along with various other salts of which comparatively small quantities are required are probably always present in sufficient amounts in the ordinary mixed diet of the average American. Calcium in the form of calcium phosphate, necessary not only for maintaining the proper supply of calcium ions in the tissues but also for building up the bony tissue of the body, is obtained largely from milk, hence the necessity for a plentiful supply of this food, particularly for growing children who are forming bony tissue.

CHAPTER VI

The Halogens

The elements chlorine (Cl), bromine (Br), and iodine (I), together with the less common fluorine (Fl), make up what is known as a *family* of elements. The different members bear a very strong resemblance to one another, the chemical and physical properties varying in regular order with increase or decrease in the atomic weight, from fluorine which is the lightest up to iodine which is the heaviest.

This family is known collectively as the *halogens* (from two Greek words meaning salt formers), so called because they are found in sea water and it is largely to compounds of these elements, particularly of chlorine, that sea water owes its salty taste. They are monovalent, and combine with hydrogen to form acids, hydrochloric acid, HCl, hydrobromic acid, HBr, and hydriodic acid, HI; and with metals to form *halides*, sodium chloride, NaCl, sodium bromide, NaBr, sodium iodide, NaI.

Chlorine is a greenish yellow gas which can easily be condensed into a liquid. It has an intensely disagreeable, suffocating odor and a very irritating action on the mucous membrane of the nose and mouth. When dissolved in water it reacts with it, giving one molecule of hydrochloric acid and one molecule of hypochlorous acid.

$$Cl_2 + H_2O = HCl + HClO.$$

Hypochlorous acid, HClO, is an unstable compound and decomposes readily, especially in the sunlight, setting free oxygen.

$$HClO = HCl + O.$$

The oxygen thus set free is in a particularly active state and therefore chlorine in the presence of moisture is a most effective oxidizing agent. It is largely used for bleaching, as it oxidizes vegetable coloring matter to colorless compounds. It also finds extensive use in the purification of water, the bacterial and other impurities being oxidized to harmless substances.

Bleaching powder is the calcium compound of the mixed hydrochloric and hypochlorous acids formed when chlorine is dissolved in water. It decomposes like hypochlorous acid itself, setting free oxygen, and is therefore an oxidizing agent. It is much used for bleaching, disinfecting, and deodorizing. Being a solid it is in more convenient form for ordinary use than the free chlorine.

Bromine is a heavy dark red liquid which gives off irritating fumes and has a very corrosive action on the skin, causing a serious burn if allowed to remain in contact with it even for a moment. The free element is not used except in the preparation of chemicals. Sodium, potassium, and ammonium bromides are all used medicinally as sedatives.

Iodine is a metallic-looking, dark grey solid which gives off violet fumes on warming. It is almost insoluble in water, but dissolves readily in alcohol and in an aqueous solution of potassium iodide. A mixture of iodine and potassium iodide dissolved in alcohol is known as *tincture* of iodine. Free iodine has valuable antiseptic properties and is much used in the dressing of wounds, etc. Various compounds of iodine are used in medicine.

CHAPTER VII

Catalysts and Enzymes.

Some chemical reactions will take place either not at all or only very slowly except in the presence of some other substance which is itself not altered in the course of the reaction. Such a substance is called a *catalyst* or *catalytic* agent, from two greek words meaning to loosen, since the action of the catalyst seems to consist in a loosening or setting free of the forces which bring about the reaction. The most striking peculiarity about catalysts is that very small quantities are required, sometimes the most minute trace of catalyst being sufficient to cause reaction between large amounts of substances, and even this small amount can be recovered unchanged when the reaction has come to an end.

The most diverse substances act as catalysts. Platinum, especially when very finely divided or spongy, is used in many commercial processes. Copper is a very efficient catalyst in certain cases. Many reactions will take place in the presence of a trace of moisture which would not go on if the reagents used were perfectly dry; for instance hydrogen and oxygen which combine with explosive violence when heated to about 1100° F. can be heated to almost twice this temperature without explosion if care has been taken to have both gases absolutely dry.

The ease and rapidity with which the most complicated and difficult reactions can be carried on in the living cell without the aid of high temperature or strong reagents could only be explained after the discovery that innumer-

able catalysts exist in plant and animal organisms. Some of these catalysts can be separated from the cells and will exert their characteristic influence on reactions outside of the organism; others have never been isolated, but their presence is recognized by the reactions which they bring about. These catalysts which are produced by the living organism and by means of which it carries on its vital processes are called *enzymes* or *ferments*. As is the case with all catalysts, a very small amount of enzyme will bring about a large amount of reaction. Each enzyme influences one special reaction and, so far as we know, only one. Even very closely related substances like the different sugars have each a special enzyme which acts upon it causing one specific kind of reaction. So in the digestive juices there is an enzyme to act on cane sugar, an enzyme to act on milk sugar, and an enzyme to act on malt sugar, decomposing these to simpler sugars which may then be acted on by still other enzymes causing further decomposition or building these up into new complexes, according to the needs of the body.

Enzymes are very sensitive to conditions. They can stand a cestain rather limited range of temperature, but there is for each enzyme a minimum temperature below which it cannot act, a maximum temperature above which it cannot act, and an optimum temperature at which its effect is most marked. Moreover they are very sensitive to small amounts of strong acids or alkali. Some enzymes will act only in faintly alkaline solution, some only in somewhat acid solution, and some prefer a neutral medium. In every case even a slight excess of either acid or alkali will destroy the enzyme.

Enzymes are frequently produced in the cell in a peculiar form known as *zymogens*. Zymogens are much more resistive to unfavorable influence of heat, acid, or

alkali than are the enzymes. They are inactive, but can be converted into the active form (activated) by other substances which the cell produces when necessary. The enzymes are classified and named according to the substances on which they act, with the exception of a few which were known and named before this system of nomenclature was devised. So the enzymes which decompose proteins are called *proteases* or *proteolytic* enzymes, but the two most important members of this class which are found in the stomach and intestines are known as pepsin and trypsin respectively. The enzymes which act on fats are called *lipases* or *lipolytic* enzymes (from *lipin*, a term used to include all fat-like substances). The starch hydrolyzing enzymes are called *amylases* (from the Latin *amylum*, starch), or *diastases* or *diastatic* enzymes from their relation to the original diastase, the starch hydrolyzing enzyme of the barley grain.

CHAPTER VIII

Nitrogen and the Atmosphere

If a lighted candle is put in an enclosed space it burns for a time, the flame becoming dimmer and more smoky, until it finally goes out. If a fresh candle or a burning match be inserted in this same space it will be immediately extinguished. Knowing that combustion is a process of oxidation we can explain this as due to the using up of the oxygen of the air by the burning candle, combustion being less and less complete as the oxygen becomes more and more exhausted, until finally there is no more oxygen and there can be no more burning. What is left when all the oxygen of the air is used up? Experiment shows that the residue is a mixture of the gas nitrogen (N) with traces of other gases such as carbon dioxide, water vapor, etc. This mixture resembles air in its appearance and superficial character but an animal placed in it dies of suffocation. If more oxygen is added to it the gas takes on once more the character of atmospheric air, without evidence of any chemical reaction taking place. From this and from the fact that the proportions of the different gases vary from time to time without altering the character of the air we conclude that the atmosphere is a mixture of gases, chiefly oxygen and nitrogen in the proportion of one part of oxygen to four of nitrogen, not a compound formed by the combination of these gases.

Animal life necessitates taking in oxygen through the lungs, using it to oxidize the complex substances of the tissues, and breathing it out again as carbon dioxide, the

result being that the oxygen of the atmosphere is decreased and the carbon dioxide increased by living animals. The following table shows the average relative composition of air as breathed into and breathed out from the lungs.

	Inspired	Expired
Oxygen	20.94	16.4
Carbon dioxide	0.03	4.1
Nitrogen	78.09	78.09

Plants, on the other hand, while using a certain amount of oxygen for this same purpose, are at the same time taking in carbon dioxide to form complex molecules which are reduction products and therefore contain less oxygen than the carbon dioxide used in their formation. This excess of oxygen is given up by the plant to the atmosphere, so that plant life results in diminishing the carbon dioxide and increasing the oxygen of the atmosphere.

In considering the cause of the ill-effects produced by "bad air," that is the air of an ill-ventilated room, the relative proportion of carbon dioxide and oxygen present was formerly regarded as of prime importance. This was a natural supposition, since the importance of oxygen for life has long been known, as has also the fact that carbon dioxide will not support life. Of late years, however, it has been demonstrated beyond possibility of doubt that the variation in these factors due to even the worst conditions of overcrowding and lack of ventilation is much too small to be of any practical importance. Experience has shown that the oxygen of the air can be reduced to less than 17 per cent., that is to a proportion which will no longer support combustion, before any ill effects can be observed. It would seem therefore that any air in which a match will burn contains sufficient oxygen for a human being to breathe. Similarly, it

has been shown that carbon dioxide does not become harmful until it reaches 1 per cent., while in one experiment a man spent twenty-four hours in an atmosphere containing over 2 per cent. and was observed to be in unusually good spirits during this time. In crowded rooms the carbon dioxide rarely reaches 0.25 per cent., so it would seem that this factor also may be disregarded.

Another theory supposed that some organic substance was present in expired air, in traces too minute to be detected by ordinary means, but very toxic even in such small amount. This view was supported to some extent by the very unpleasant odor of a room crowded with human beings. Absolutely no proof has been found, however, of any such toxic substance. On the contrary, there is evidence to show that even the most vitiated air can be breathed into the lungs without injury. A number of people were confined in a small air-tight room until they began to grow faint and ill, while another man, stationed outside the room but breathing the same air through a tube felt perfectly comfortable. On the other hand no relief was felt by the sufferers when they were allowed to breathe the outside air through a tube. Then an electric fan was started in the room without changing the air in any other way than by setting it in motion. Those inside immediately revived and felt as though fresh air had been admitted. From this and many other similar experiements hygienists have come to the conclusion that it is not the chemical composition of the air which is important but the temperature and amount of moisture present.

The human body is continually producing heat, most of which passes out from the body through the skin, either by radiation into the atmosphere or by evaporation of the moisture on the surface of the skin. If its escape is prevented the temperature of the body rises, the nor-

mal action of the cells is disturbed, lassitude and fatigue are felt, followed by all the symptoms of heat prostration. Accumulation of heat in the body is always unfavorable and may be caused either by high temperature of the surrounding air which prevents radiation of heat from the body, or by high humidity (moisture content) which prevents rapid evaporation of moisture from the skin. It is well known that a warm humid day in summer causes more heat prostrations than a much hotter day when the air is drier. The ill-effects of humidity in our rooms in winter are not so obvious, but make themselves felt none the less. Unfortunately the ideal humidity at different temperatures has not yet been ascertained. Too little moisture makes an uncomfortable atmosphere just as does too much, and this is perhaps a more common defect in our artificially heated houses. Over-heating is equally or perhaps more common amongst us. A body will radiate heat only when it is in contact with something cooler than itself, and the greater the difference in temperature the more rapidly does radiation take place. The higher the temperature of the air therefore the less heat is lost by radiation from the body. If the air is dry this may be partially compensated by the increased evaporation, but if the air is both moist and hot both methods of escape are prevented and the heat rapidly accumulates to a dangerous degree. Hygienists agree that except for old people, who need a slightly higher temperature, a room should not be kept above 70°. There is a little more difference of opinion about the desirable lower limit. In this country it is usually put about 65°, but the English authorities consider 60°, or even below that, more healthful. Probably custom has a good deal to do with it, as the body is known to have considerable power of adaptation and can gradually be accustomed to a temperature many degrees lower than

that previously found comfortable. It must be remembered here, however, as in all cases where a change in the living conditions of a human being is concerned, that the change must not be made too suddenly. Sufficient time must always be allowed for readjustment to take place if discomfort is to be avoided.

Occasional changes of temperature, within reason, seem better than a steady level, as the change has a stimulating effect. Also, air that is in motion is better than air which is perfectly still as the currents convey the hot moist air away from the body and thus assist both radiation and evaporation. The ill effects of drafts on the other hand seem to be due to the chilling down of the body at one point too suddenly to allow of its protecting itself by producing more heat as it does when the reduction is more gradual or more general. Just when moving air ceases to be beneficial and becomes an injurious draft is a little difficult to determine. A good deal depends on the individual, as some people are more sensitive to changes of temperature than are others. Also the state of health of the individual is important, as a robust person can withstand a shock under which a weaker person would collapse.

Although we have entirely given up the idea that it is necessary to keep our windows open in order to secure an adequate supply of oxygen for our lungs we are still largely dependent on window ventilation to cool and humidify our air and keep it in motion. Especially at night, or in a sick-room where the patient is warmly covered up in bed and protected from chill, it is possible to regulate the temperature very satisfactorily by this means. Doubtless in the course of time our engineers and architects will co-operate to produce buildings in which the ideal conditions will be secured automatically, but meantime probably the best advice we can give

the public in general, especially our heat-loving American public, is Keep your windows open. It must be remembered, however, that the purpose of the open window is chiefly to prevent over-heating, and that this may be brought about also by economy of fuel, a point which is of some importance to families of limited income or in time of a coal famine. It is perhaps fortunate for the public health that the majority of modern houses have such loose-fitting doors and windows as to secure ventilation with or without the will of the occupants.

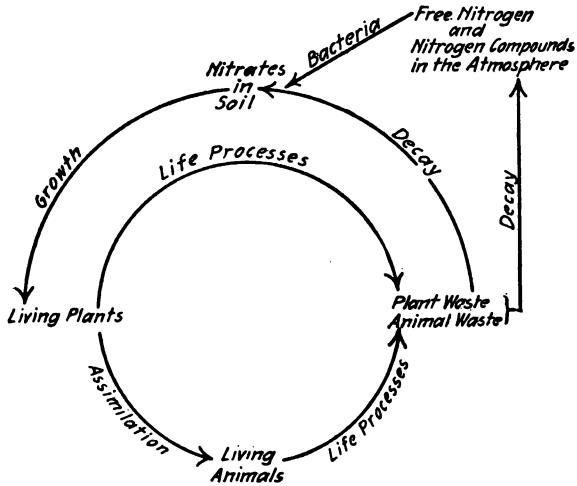
CHAPTER IX

Compounds of Nitrogen

Free, that is, uncombined nitrogen is comparatively uninteresting, but its compounds are of fundamental importance. Organic nitrogen derivatives known as proteins (see page 104) are used in building up the muscular tissue of animals, and it is therefore necessary that a certain amount of this element in a suitable form be supplied regularly in the food. Animals have no power of making use of free nitrogen and the enormous supply of nitrogen in the atmosphere is therefore unavailable for them. Neither can they use the inorganic salts containing nitrogen, such as the nitrates and nitrites (see page 37) present in the soil. They are therefore dependent on the organic nitrogen compounds which they can obtain from the plants, either directly or, in the case of flesh eating animals, after previous consumption and assimilation by some other animal. The plants obtain their nitrogen from nitrates of the soil which are soluble and can therefore be drawn up through the roots in solution and are then combined with carbon and hydrogen to form plant protein. The growth of plants results therefore in the gradual removal of nitrates from the soil, and these must be restored to it if it is to remain a good medium for plant life. This is sometimes accomplished by the use of artificial fertilizers containing nitrates, or by natural fertilizers such as animal excreta or dead animal or vegetable matter which in decaying give up their nitrogen and return at least part of it to the soil again in soluble form. Although green plants, like animals, are unable to use the nitrogen supply of

the air certain forms of bacteria have the power of converting this free nitrogen directly into the useful nitrates. These bacteria grow in little clusters on the roots of leguminous plants, peas, beans, clover, etc. supplying more nitrates than the plant is able to use in its lifetime, so that the growth of such plants results in the enriching of the soil instead of exhausting it as do other plants. In recent years successful commercial methods have been devised for the *fixation of nitrogen*, as it is called, that is, the conversion of the atmospheric nitrogen into compounds which can be used as plant food. These are of the utmost practical importance as they insure an inexhaustible supply of the fertilizers which are directly essential for vegetation and indirectly for the maintenance of animal life on the earth.

The cycle through which nitrogen passes in nature may be represented by the following diagram:



There are two general types of inorganic nitrogen compounds; ammonia, NH_3, and its derivatives, and the acids containing nitrogen and their derivatives.

Ammonia is a pungent smelling gas which is soluble in water, combining with it to form a base, *ammonium hydroxide.*

$$NH_3 + H_2O = NH_4OH$$

On warming, the ammonium hydroxide breaks up readily into NH_3 and H_2O again. The *household ammonia* used in cleaning is a dilute solution of ammonium hydroxide. Both ammonia gas and ammonium hydroxide react with acids to form *ammonium salts.*

$$NH_3 + HCl = NH_4Cl$$

$$NH_4OH + HCl = NH_4Cl + H_2O$$

Many of these salts have the same tendency towards decomposition as ammonium hydroxide and hence have a strong odor of ammonia gas.

$$NH_4Cl = NH_3 + HCl$$

Smelling salts are made by mixing such a salt of ammonia with some pleasant-smelling flower extract like oil of lavendar and keeping the mixture tightly corked. The odor given off when the stopper is removed has a refreshing and stimulating effect.

Ammonia gas is quite easily condensed to a liquid by applying pressure. When the pressure is removed the liquid returns to the gaseous state again. The difference between a substance as liquid and the same substance in the gaseous state consists in the different amount of space between the molecules, the molecules of a liquid being held close together while the molecules of a gas are so far separate that each individual molecule has very little influence on its neighbors. To separate the particles of a liquid requires work, or expenditure of energy. In our common experience this energy is supplied in the form of heat, as when we boil water to convert it into steam.

Boiling water never becomes hotter than 180°F., no matter how much fire we put under it, because the heat is used up in separating the molecules from one another and so converting the liquid water into the gaseous steam. The more heat, the faster this separation goes on, but the temperature will not rise so long as there is any liquid water left. When any liquid *evaporates*, that is passes from liquid to vapor, heat must be used up, and it takes this heat from the surrounding air and consequently cools it down, but if the evaporation is slow this loss of heat is so gradual as to be unnoticeable. The more volatile the liquid, that is, the more rapidly it tends to pass into vapor, the more pronounced is the change in temperature. Evaporation of moisture from the skin gradually absorbs heat from the blood and reduces the body temperature, hence the benefit of perspiration in preventing dangerous overheating of the system. If a drop of ether is placed on the skin it evaporates a great deal more rapidly than water, taking so much heat from the skin as it does so that a sensation of intense cold is produced. When liquid ammonia evaporates it does so very quickly, absorbing much heat and consequently lowering the temperature of its surroundings to a considerable degree. If pipes through which liquid ammonia is circulating are surrounded with water so much heat will be taken from the water that it freezes into solid ice. It is upon this principle that the use of ammonia in refrigerating plants and in making artificial ice depends. Ammonia gas is compressed into liquid and circulated through a system of pipes where it absorbs heat and evaporates, only to be compressed once more and enter into circulation again, making a continuous process.

Nitric acid, HNO_3, is the most important acid containing nitrogen. Besides being a strong acid it is an

excellent oxidizing agent, owing to its tendency to break up and set free oxygen.

$$2HNO_3 = H_2O + 2NO_2 + O$$

This property is shared by many derivatives of nitric acid, especially its organic compounds. Our most powerful explosives belong in this class, and owe their destructive power to the fact that they decompose with great readiness giving gaseous products which expand into infinitely greater volume than the substance from which they were formed, and in this sudden expansion clear everything from before them.

CHAPTER X

The Metals

Those elements which can take the place of hydrogen in an acid, forming salts, are called *metals*. The metals in general are solids, although one, mercury, is a liquid. They are good conductors of heat and of electricity, and have a bright lustre which is known as a "metallic" lustre. Although there is a large number of metallic elements only a few are in common use. These are platinum, gold, silver, iron, copper, aluminum, tin, nickle, lead, and zinc.

Platinum, Pt, and *gold*, Au, are sometimes called the "noble" metals, because they are not acted upon by ordinary reagents. The common acids will not dissolve them, nor do they tarnish in the air as do most other metals. Platinum is an excellent conductor of electricity and has a high melting point. Like iron, it softens before melting, which makes it possible to weld it. It can be hammered into sheets or drawn out into wires, and, as it expands and contracts with change of temperature at about the same rate that glass does, it can be fused into glass. These properties, along with its resistance to chemical agents, make it very valuable for the construction of scientific instruments. It is comparatively soft, however, and can be rather easily bent or scratched.

Silver, Ag, resembles platinum in appearance and certain physical properties, but is much more easily acted on by reagents. It dissolves readily in nitric acid, forming silver nitrate, the substance sometimes known as

"lunar caustic" and used in medicine as a disinfectant and for cauterizing. The oxygen of the air has no effect on silver, but if the atmosphere is contaminated by traces of sulphur compounds, as is apt to be the case where soft coal or illuminating gas is used, the silver turns black, owing to the formation of a film of silver sulphide. This is what is known as the tarnishing of silver. The blackening of silver spoons when used for eggs is due to the sulphur in the egg. Wool contains sulphur, as does also vulcanized rubber, and therefore silver tarnishes quickly if left in contact with either of these.

The polishing of silver consists in removing this surface of sulphide. This can be done by rubbing with some substance which is harder than the silver but so finely powdered that it will not scratch the metal. Rouge (ferric oxide), whiting (powdered chalk), and infusorial earth (made up of the shells of microscopic animals) are some of the best substances for this purpose. Sand is sometimes found in polishing powders, but is not a desirable ingredient, owing to the difficulty of reducing it to a sufficiently fine powder. These various powders are sometimes made up into a paste for convenience in handling, but they have little, if any, advantage over the powders. In either case vigorous rubbing is necessary to produce a polish. Other silver polishes contain some chemical which will dissolve silver sulphide. With these it is not necessary to rub hard in order to remove the tarnish, but silver thus cleaned has a dull appearance unless burnished by rubbing it briskly with chamois leather or a soft cloth after the polish has been removed. The silver polishing pans on the market are pans of aluminum or zinc in which the silver is boiled with a solution of washing soda and salt. By a series of reactions which are electro-chemical in their nature, the silver sulphide is decomposed and the sulphur removed.

Here again the silver is left clean but not bright and requires burnishing.

Certain silver plating preparations are on the market, to be used instead of a polish. These contain a small amount of silver in the form of a compound from which it is deposited in a thin film when it comes in contact with the article to be polished. This is an advantage in the case of thinly plated silver, which tends to wear off in the course of time.

Pure silver is too soft for ordinary use, and is hardened by addition of a certain amount of copper. "Sterling" silver contains 92.5 per cent. silver and 7.5 per cent. copper. On account of the copper present silver utensils should never be left in contact with vinegar or with foods containing vinegar, as the acetic acid of the vinegar will react with the copper forming a poisonous copper salt.

Iron, Fe, is the most widely used of all metals. Three varieties of iron are known in the industries; cast iron, steel, and wrought iron. Wrought iron is almost pure iron, but cast iron and steel contain varying small amounts of other substances, of which the most important is carbon. It is to the proportions of carbon present that these two varieties owe their specific properties. Cast iron is rather brittle, whereas steel is tougher and can be tempered to almost any desired degree of hardness and elasticity. Overheating or too sudden cooling from a high temperature will change the temper of steel, and for this reason steel implements should never be subjected to such treatment.

The rusting of iron is a process of oxidation which takes place only in the presence of moist air. In perfectly dry air no rust appears, nor will iron rust under water free from dissolved air. All iron or steel utensils should therefore be very carefully dried and kept from damp

ness. Moreover, any trace of rust should be immediately cleaned off, as its presence causes the rusting process to go on more rapidly.

Various materials are used to coat iron and protect it from the air. Tin is generally used for this purpose in making kitchen utensils since tin it self is not affected by air or moisture. If the tin layer is scratched, however, a little electric current is set up between the tin and the iron underneath, and the iron rusts faster than it would if there were no tin there.

If iron is treated with superheated steam at a high temperature it forms another oxide with slightly different proportions of iron and oxygen from those in ordinary iron rust. This oxide, instead of scaling off and leaving fresh surface exposed to the action of the air as rust does, adheres firmly to the surface of the iron and protects it from any further action. Roasting pans and bread pans are made of iron so treated, and are therefore quite resistive to rust, provided they are kept dry when not in use.

Iron covered with zinc is called *galvanized* iron. Zinc is a more effective covering than tin, because even though the coating should become scratched or worn through the presence of zinc hinders the oxidation of the iron instead of promoting it as does tin. Galvanized iron cannot be used for cooking utensils, as zinc forms poisonous salts with vegetable acids, but it is very satisfactory for water pipes, pails, and so on.

The *enamelled ware* so much used is iron covered with a glaze similar to that used for china. It is quite impervious to weak alkalies and to the dilute acids of fruits and vegetables, and if properly put on the glaze will not chip off with ordinary handling. Dry enamel ware should not be subjected to sudden heating or cooling however,

as even the best glaze is liable to crack under those conditions.

Iron for water taps and fixtures of various kinds is sometimes coated with *nickel.* Nickel has very little tendency to tarnish and is easily kept bright with soap and water, with occasional rubbing with kerosene oil or scouring to remove the film of grease and dirt which accumulates.

Copper, Cu, is said to be, next to iron, the most useful metal. Enormous quantities are used for wiring in electrical industries, for roofing, for nails, rivets and sheathing in ship building, and so on. In the presence of moist air copper tarnishes rapidly, becoming coated either with brown oxide or with a green carbonate known as verdigris, formed from the carbon dioxide in the air. Copper itself is not acted on by organic acids, so bright copper utensils can be used in cooking with safety. If they are allowed to become tarnished, however, the organic acids will act on the oxide or carbonate as the case may be, forming poisonous salts. It is most essential therefore that all such utensils should be kept in a condition of high polish. Even polished copper vessels are apt to contaminate the food with small traces of copper, since almost all fats, oils, and syrups dissolve copper to some extent, but the amount is so small that it is not believed to be injurious to health.

Green vegetables cooked in copper boilers retain their green color much better than those cooked in vessels of iron, tin, or enamel. This is sometimes made use of by canners to make their products look attractive, and it used to be not uncommon to find canned peas to which so much copper had been added that they left a distinctly metallic taste in the mouth. Such highly contaminated vegetables are undoubtedly injurious, but fortunately the public has been taught to look with suspicion on

too highly colored foods, and this bad practice is no longer very common.

Ammonia acts on copper in the presence of air, forming a deep blue solution. Ammonia is sometimes used in polishes for copper and brass. When such a polish is used all traces of it should be carefully removed, as its presence would promote further oxidation of the metal. The same thing is true of polishes containing weak acids.

Aluminium, Al, oxidizes superficially in moist air, but the oxide, which is white, adheres firmly to the surface of the metal, and can be rubbed to a polish almost as bright as that of the metal itself. Since in addition to its resistance to air it is very light, it makes an excellent material for instruments and utensils of all kinds. Dilute organic acids have little, if any, solvent action on aluminium, but the presence of sodium chloride is said to increase the action. The amount of soluble aluminium salts formed is however probably inconsiderable under any ordinary conditions. Salt solutions in general (*e.g.* sea-water) corrode aluminium rapidly, even the small amount of salts present in hard water having a corrosive action. It is dissolved rapidly by alkalies, hence these should never be used in cleaning aluminium. Special scouring powders and pastes free from alkalies may be obtained for the purpose, but as a general thing it may be kept bright by rubbing dry after each washing and occasionally wiping with a little kerosene oil. If an ordinary metal polish is used care should be taken to remove every trace of it from the metal.

Lead, Pb, is largely used in the arts on account of the ease with which it can be worked and soldered. It tarnishes slightly, but only superficially, in moist air. The organic acids of fruits and vegetables act upon it, forming soluble salts, which like practically all compounds of lead, are poisonous. Lead is an ingredient of

common solder, and therefore soldered utensils should never be used for cooking foods which may contain acids. Tin cans used to be plated with tin containing an appreciable amount of lead, and were sealed with a lead solder. As long as the cans were air tight the acids present had no effect on the lead, but when the cans were opened and the oxygen of the air had free access lead salts went into solution. For this reason it was unsafe to leave food standing in an open can. Modern methods of making and sealing cans have lessened the danger, but it is still wiser to take this simple precaution as a general rule.

Lead has been used for water pipes since the days of the Romans. Lead pipe cannot be used for soft water, since this, by means of the oxygen and carbon dioxide which it has dissolved from the air, reacts upon the lead, forming compounds which dissolve or diffuse through the water rendering it unfit for drinking. Continued use of water containing as little as 0.00005 per cent., of lead is unsafe, since lead compounds are excreted from the system with difficulty and tend to accumulate to a serious extent. Hard water has in general less effect on lead, and in some cases where the water supply of a town was found to be acting on the lead pipes the difficulty has been overcome by filtering the water through a bed of limestone or chalk to supply the necessary hardness. The custom which prevails in many parts of the country of using lead-lined cisterns for the storage of rain water is a dangerous one and may result in serious contamination of the water. Where lead lined tanks must be used they should be well painted on the inside with a carbon paint and should be frequently examined for defects in the coating.

Alloys are mixtures of two or more metals which have been melted together and allowed to solidify into a solid

mass. The physical properties of an alloy, such as color, hardness, and melting point, are different in many cases from those of the separate ingredients, but the individual metals retain their chemical properties to a large extent. In general an alloy will melt lower than a pure metal, and alloys made by combining together various metals which themselves have rather low melting points will melt in warm water. Low-melting alloys, or fusible metals, as they are called, are used in making fuses for electrical connections. If too strong a current is led through the wires the heat developed melts the alloy and the circuit is broken before any serious damage has been done. Fusible metal is also used as plugs in automatic sprinklers. Should a fire start at night in an empty building the heat melts the plugs of the sprinklers and a deluge of water descends into the room below.

Copper is an ingredient of many alloys. Brass is an alloy of zinc and copper; bronze an alloy of copper, zinc, and tin; gun metal is tin and copper; "german silver" contains copper, zinc, and nickel. In general, these alloys should be treated with the same care as copper itself, although the alloys are much less liable to tarnishing than is the pure metal.

SECTION II

ORGANIC CHEMISTRY

CHAPTER XI

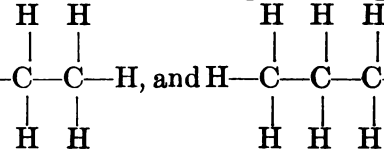
Hydrocarbons

The term *organic* chemistry was originally confined to the chemistry of compounds which were produced by the vital activity of either animals or plants, on the supposition that such compounds differed essentially from those which were produced without the agency of the life force. When first one and then another of these vital products had been reproduced in the laboratory it was recognized that this distinction must be done away with, but the name "organic" was retained for the vast number of compounds of carbon and hydrogen, with or without other elements present, on account of the definite relations which were found to exist between these substances.

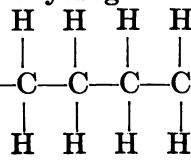
The basis of organic chemistry is the tetravalent carbon atom, which can be pictured thus, $-\overset{|}{\underset{|}{C}}-$, the four lines representing four valences which are joined to other atoms. The simplest organic compound is that in which each of these valences is joined to a hydrogen atom. This is the gas *methane*, the formula of which may be written $H-\overset{H}{\underset{H}{C}}-H$, or, more simply, CH_4. The

former method of representation is what is known as a *graphic* or *structural* formula, that is, a formula so written as to show the exact arrangement of the four valences. In a compound so simple as methane this is of no consequence, but in more complex molecules it is of the greatest importance. Organic chemistry differs materially from inorganic in this particular, that whereas in inorganic chemistry a formula stands for one definite compound, among the organic compounds a very large number may be represented by one single formula. This is due to the great variety of combinations possible where several carbon atoms, each with four valences, are linked together. This may be illustrated by the formula C_4H_{10}. This formula represents a compound containing four carbon atoms and ten hydrogen atoms in the molecule, but experience shows that there are actually two compounds with distinct properties each having this composition. On the other hand, there is only one CH_4, one C_2H_6, and one C_3H_8.

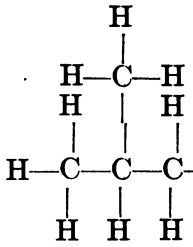
A study of the structural formulæ enables us to see how this may be possible. There is only one way in which one tetravalent carbon atom and four monovalent hydrogen atoms can be joined up together; that is as we have written it. Nor can we arrange two carbons and six hydrogens, or three carbons and eight hydrogens in any



but one way, H—C—C—H, and H—C—C—C—H. But four carbons and ten hydrogens can be linked together in



one long chain, H—C—C—C—C—H, or the chain may

be branched  There are the same number of atoms in each case, but they are represented as occupying different relative positions in space. Compounds like these, which have the same composition but distinct properties and which are supposed to differ only in the relative positions of the atoms in the molecule are called *isomers*, or *isomeric compounds*. The number of possible isomers increases very rapidly with increasing number of carbon atoms or when other atoms or groups of atoms are substituted for one or more of the hydrogens in the molecule. Although there is only one C_3H_8, there are two C_3H_7I's, namely, $CH_3.CH_2.CH_2I$, and $CH_3.CHI.CH_3$.

Compounds containing carbon and hydrogen only are grouped together into one general class, the *hydrocarbons*. A hydrocarbon in which every carbon valence is attached to a separate atom or group of atoms, as in the formulæ given above is said to be *saturated*. In an *unsaturated* hydrocarbon two carbon atoms are linked together by two, or sometimes by three valences, as in ethylene, $CH_2=CH_2$, and acetylene, $CH\equiv CH$. A carbon atom which has more than one of its valences going to any other atom is said to be unsaturated. Unsaturation is always characterized by great reactivity. The *double* and *triple bonds*, as they are called, readily separate, adding on other atoms to form a saturated compound. The saturated compounds cannot undergo addition reactions, and react only by substitution, by

exchanging one of the atoms already present in the molecule for some other atom.

Paraffins are hydrocarbons composed of open chains of saturated carbon atoms. Chemically they are very inert, reacting only with the halogens, Cl, Br, and I, with which they give halogen substitution products in which one, two, or more hydrogens are replaced by halogen atoms. Taking the reaction of methane with bromine as typical, under suitable conditions the reaction proceeds as follows:

$$CH_4 + Br_2 = CH_3Br + HBr.$$

In presence of excess of bromine the reaction may be carried further until all four H's have been successively replaced.

$$CH_3Br + Br_2 = CH_2Br_2 + HBr.$$

$$CH_2Br_2 + Br_2 = CHBr_3 + HBr.$$

$$CHBr_3 + Br_2 = CBr_4 + HBr.$$

Some of these halogen substitution products are of practical importance. $CHCl_3$ is the valuable anæsthetic *chloroform;* the corresponding iodine compound, CHI_3, is an antiseptic, *iodoform;* CCl_4, *carbon tetrachloride* is an excellent solvent for fats and has found considerable use as a cleansing agent, being sold for this purpose under the trade name *carbona.*

Although the group of paraffins includes various substances in common use, such as *gasoline, kerosene, benzine,* and other cleansing fluids, and many lubricating oils, their chief interest from the chemist's point of view lies in the fact that they are the parent substances of an almost infinite number of derivatives, to which is given the general name of *aliphatic compounds.* While the paraffins do not react directly with anything except the

halogens, the halogen derivatives thus formed are more reactive, and it is possible to substitute for the halogen almost any desired atom or group of atoms. In this way are derived a great many different classes of aliphatic compounds, many of them somewhat resembling the various types of inorganic compounds, oxides, hydroxides, acids, and so on, but differing essentially from these in many of their properties and reactions. Only a few of these classes, and only one or two of the more important members of each class will be considered here. It is to be remembered that, with the exception of a few minor peculiarities, what is true of one individual substance is true of the whole class to which it belongs. It is therefore only necessary to learn a few general facts which may easily be kept in mind, without burdening the memory with a mass of detail.

The residue which would be left if we took one hydrogen from a paraffin is called an *alkyl group*, or *alkyl radicle*. An alkyl group cannot exist by itself since it would have a valence unsatisfied, but all aliphatic compounds except the paraffins themselves may be considered alkyl derivatives. In the various reactions in which these derivatives take part the alkyl group remains for the most part unchanged, behaving like a single unit. Halogens are the only reagents which affect the alkyl group, and they may be made to react with it as they do with the paraffins themselves, giving halogen-substituted groups. The individual alkyl groups are named as follows: CH_3, *methyl;* C_2H_5, *ethyl;* C_3H_7, *propyl;* C_4H_9, *butyl;* C_5H_{11}, *amyl;* and all others by adding the suffix *-yl* to the stem of the Greek numeral indicating the number of carbon atoms which they contain, *hexyl, heptyl, octyl,* etc. Addition of one hydrogen to an alkyl group gives the corresponding paraffin, which is named by substituting the ending

-ane for the *-yl* of the alkyl group; thus CH_4 is *methane*, C_2H_6, *ethane*, C_3H_8, *propane*, etc.

It is to be noted that the various alkyl groups, and, similarly, the various paraffins, differ from each other in composition only by addition or subtraction of CH_2 or a multiple of this. Any series of compounds in which the members differ by this regular progression is called a *homologous series*. The members of a homologous series resemble each other very closely in chemical properties, differing only in degree of reactivity, the lower members, *i.e.* those with fewer carbon atoms, being in general slightly more reactive than the higher members. In physical properties a regular gradation may be observed, the lower members being as a rule more volatile and more soluble than the higher. The lower members of any series are apt to be gases or liquids, the higher members solid.

CHAPTER XII

Alcohols and Ethers

If one hydrogen in a paraffin is replaced by an hydroxyl group the resulting compound is an *alcohol.* The alcohol molecule may therefore be regarded as composed of an alkyl group and an hydroxyl group, the latter giving to the molecule its characteristic properties. The simplest member of the series of alcohols would be CH_3OH; the next higher member, according to our definition of homologous series, would be C_2H_5OH, the next C_3H_7OH, etc. These compounds are named from the alkyl group which they contain, methyl, ethyl, propyl, etc., alcohols. An alcohol containing only one hydroxyl group in the molecule is called *monohydric;* those containing two or more hydroxyls are called *polyhydric.* The most important polyhydric alcohol is *glycerine,* $C_3H_5(OH)_3$, derived from propane by replacing one of the H's attached to each C by OH; so from $CH_3.CH_2.CH_3$ is derived $CH_2OH.CHOH.CH_2OH$.

In general the alcohols are colorless liquids, although a few of the higher members are solid. The liquid members are good solvents for almost all kinds of organic compounds, and for that reason play an important part in a great number of very diverse organic reactions. Moreover, in virtue of the reactivity of the hydroxyl group they are used in many reactions for the purpose of introducing either an alkyl or an alkoxyl ($-OCH_3$, $-OC_2H_5$, etc.) group into a molecule.

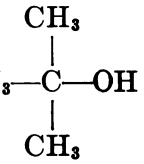
Methyl alcohol, CH_3OH, is sometimes called *wood spirit* because it is one of the principle products obtained

when wood is heated to a high temperature in a retort out of contact with air. Because of this very cheap method of production methyl alcohol is the cheapest of the alcohols, and since it is as good a solvent for most purposes as the more costly ethyl alcohol it was formerly used very extensively in manufactures. Unfortunately it is extremely poisonous, acting not only when taken into the system through the mouth, but even when absorbed through the skin or the lungs. Workmen engaged in the manufactures in which methyl alcohol was extensively used suffered severely from general debility, followed in many cases by partial or even complete blindness and paralysis. Pure ethyl alcohol on the other hand is too expensive for most industrial uses, as in addition to the cost of making there is a heavy Government tax imposed on all alcohol which can be used as a beverage. *Denatured alcohol* is ethyl alcohol to which has been added some other substance with a disagreeable taste and smell but which will not affect its use for industrial purposes. Such alcohol is tax free, and sufficiently cheap to enable the manufacturer to substitute it in almost every case for the disagreeable and dangerous methyl alcohol. The first denaturant used was 10 per cent. methyl alcohol, the mixture being sold as *methylated spirit;* later an alternative denaturant, 2 per cent. methyl alcohol and 0.5 per cent. of certain pyridine bases, was permitted by the Government. Since 1917 alcohol for external medical use has been denatured by the addition of a small amount of carbolic acid. This is a poison when taken internally, but was formerly much used as an antiseptic, although it has fallen somewhat into disrepute of late years.

Ethyl alcohol, the alcohol of common usage, is prepared by fermentation from the starch present in grain and potatoes or the sugar in fruit juices, a process to

be described later. When denatured it is used as a solvent, as a reagent for industrial purposes, and to some extent as a fuel. When diluted with water and flavoring extracts it is used as a beverage. *Whiskey* is grain alcohol flavored by the extractives from the grain, and contains about 40 per cent. alcohol. *Brandy* is the spirit obtained by distillation of wine, and usually contains from 40 to 50 per cent. alcohol. *Rum* is made by distilling the product obtained by fermenting molasses, and contains about 50 per cent. alcohol. *Wines* are made from fruit juices, and contain from 7 to 16 per cent. alcohol. The various *beers* are made from barley and contain, along with from 2 to 4 per cent. of alcohol, various extractives from the barley grain, to which they doubtless owe their tonic properties.

In *propyl alcohol* we find the possibility of isomerism among the alcohols. There are two possible propyl alcohols, one of which has the hydroxyl at the end of the chain, $CH_3.CH_2.CH_2OH$, and one in which it is in the middle, $CH_3.CHOH.CH_3$. An alcohol in which the hydroxyl is attached to a carbon atom which is directly joined to only one other carbon as in the first of the two formulæ given above, is called a *primary* alcohol; one in which the hydrogen is attached to a carbon atom which is directly joined to two other carbons, as in the second formula, is called a secondary alcohol; while one in which the hydroxyl is attached to a carbon which is directly joined to three other carbons is called a



tertiary alcohol, *e.g.* tertiary butyl alcohol, $CH_3—C(CH_3)_2—OH$

Ethers.—As the alcohols with their OH group correspond more or less to inorganic bases, so the *ethers* may

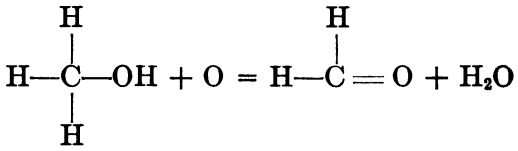
be considered analogous to inorganic oxides. The general formula of the ethers may be written R—O—R_1, where R and R_1 represent alkyl groups. They are derived from the alcohols by substituting a second alkyl group for the H of the OH. The anæsthetic known as *ether*, or sometimes *sulphuric ether* (from the fact that sulphuric acid is used in its preparation), is ethyl ether, C_2H_5—O—C_2H_5.

Ether is an excellent solvent for most organic substances. It mixes readily with alcohol in all proportions but is almost insoluble in water. When it is necessary to cleanse the hands thoroughly for surgical work they are washed first with soap and water to take off the dirt particles, then with alcohol, which mixes with and removes the last traces of moisture and soap, and then finally with ether to dissolve away the closely adhering film of grease which is present on all normal skin. Were it not for the intermediate treatment with alcohol the ether would not be able to penetrate into the pores and crevices of the skin, while alcohol alone would be a much less efficient cleansing agent.

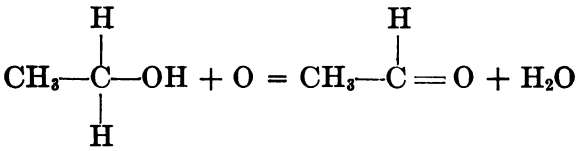
CHAPTER XIII

Aldehydes and Ketones

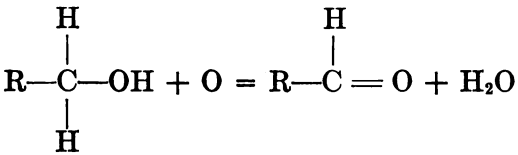
Alcohols are fairly susceptible to oxidizing agents, the lower members being the most easily oxidized. The product of oxidation depends upon the position of the hydroxyl group. Oxidation of a primary alcohol produces an *aldehyde*, with loss of two atoms of hydrogen which are oxidized to water. In the case of methyl alcohol the reaction may be written thus:



Similarly in the case of ethyl alcohol,



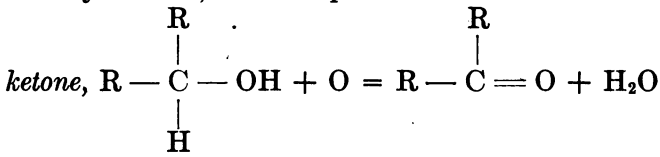
or, in general, if we use R to represent any alkyl group,



The aldehydes are characterized by the group $-C(=O)H$, the fourth valence of the carbon being satisfied by any

alkyl group, or in the case of the aldehyde derived from methyl alcohol, by another hydrogen.

When a secondary alcohol is oxidized the reaction is exactly similar, but the product in this case is called a



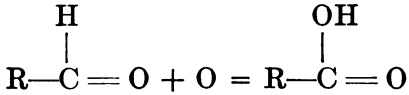
ketone, $R-\underset{\underset{H}{|}}{\overset{R}{\overset{|}{C}}}-OH + O = R-\overset{R}{\overset{|}{C}}=O + H_2O$

Ketones are characterized by the group $>C=O$, both remaining valences of the carbon being satisfied by alkyl groups.

Tertiary alcohols are oxidized only with great difficulty, and the process results in complete rupture of the molecule.

It will be noted that the grouping $>C=O$ is common to both aldehydes and ketones, the difference being that in the case of the aldehydes at least one of the remaining valences is satisfied by hydrogen, while in the ketones they are both attached to alkyl groups. This $>C=O$ grouping is called the *carbonyl* group, and is characterized by great reactivity, which is increased by the presence of hydrogen attached to the carbon of the carbonyl group. For this reason, while aldehydes and ketones are both reactive the aldehydes are much more so. For the most part they undergo the same reactions, but in the case of the aldehydes the reaction is more easily brought about

Aldehydes are very readily oxidized further, the H being oxidized to an OH.



$R-\overset{H}{\overset{|}{C}}=O + O = R-\overset{OH}{\overset{|}{C}}=O$

On account of the great ease with which this reaction takes place aldehydes are good reducing agents. Certain

metallic oxides will give up their oxygen to the aldenydes, themselves becoming reduced to the metal. This has become the basis of a simple test for an aldehyde, copper being the metal most commonly used for the purpose. In the presence of an aldehyde a solution of cupric oxide is reduced to cuprous oxide, which forms a dark red precipitate.

$$R{-}\overset{H}{\overset{|}{C}} = O + 2CuO = R{-}\overset{OH}{\overset{|}{C}} = O + Cu_2O$$

Since ketones have no hydrogen attached to the carbonyl carbon, oxidation can take place only by rupture of the chain. Thus from the simplest ketone, acetone, $CH_3.CO.CH_3$, we obtain on oxidation $CH_3{-}\overset{OH}{\overset{|}{C}} = O$, CO_2, and H_2O. As might be expected, ketones are less readily oxidized than are the aldehydes.

Formaldehyde, H.CHO, the first member of the aldehyde series, is produced by the oxidation of methyl alcohol. It is a gas with a very pungent odor, which is largely used as a disinfectant. A 40 per cent. solution of the gas in water is sold under the trade-name *Formalin*. Formaldehyde molecules have a tendency to combine together into a complex molecule of which the exact structure is unknown. The formaldehyde candles which are used in disinfecting are composed of this formaldehyde complex, or *paraformaldehyde*, as it is called. When they are lighted the heat from the burning wick causes the disintegration of the complex into the simple form. Formaldehyde acts on micro-organisms by combining with their protoplasm and preventing it from functioning.

CHAPTER XIV

Acids

Organic acids, like inorganic, owe their acid properties to hydrogen ions, but not all the hydrogen of organic acids is ionized. The H of the -OH group as found in the alcohols is non-ionizable, as is the H attached to the —CO group in the aldehydes. When an OH and a doubly bound O are attached to the same C however, as is seen in the oxidation products of the aldehydes, the H becomes ionizable and the compound shows all the properties of an acid. The grouping $-\overset{OH}{\overset{|}{C}}=O$, (or —COOH as it is oftener written), is called the *carboxyl group*, and is characteristic of the organic acids.

Acids are produced as the final oxidation products of the alcohols, aldehydes, or ketones. Those in most common use, however, are more frequently obtained as products of the vital activity of animals or plants. The souring of milk, wine, bread, and many other food substances, is due to the production of acid through the agency of bacteria. *Acetic acid*, $CH_3.COOH$, is obtained by oxidation of ethyl alcohol, the corresponding aldehyde being an intermediate product. This oxidation may be brought about by ordinary chemical means or by the action of the bacillus *Mycoderma aceti*. Since this bacillus is very commonly present in the air, on the skins of fruit, etc. acetic acid is produced whenever a dilute solution of alcohol, along with some nitrogenous material to serve as food for the micro-organisms, is

left exposed to the air. If there is more than about 10 per cent. of alcohol present however the bacteria will not flourish and the souring is prevented. Ordinary vinegar is a solution of acetic acid produced in this way, containing 3–6 per cent. of acid together with flavoring material from the fruit or grain extract used for fermenting. The name *vinegar*, without qualification, is used in the United States for the product made from apple juice or cider only, but vinegar may also be made from beer (malt vinegar), wine (white wine vinegar), or molasses. Acetic acid is also obtained, along with methyl alcohol, when wood is destructively distilled, and it is in this way that most of the acetic acid used industrially is obtained.

The acid produced when milk sours is *lactic acid*, $CH_3.CHOH.COOH$. This acid has a hydroxyl group in the molecule as well as a carboxyl, and therefore shows some of the properties of an alcohol as well as those of an acid. *Tartaric acid*, present in the form of its sodium-potassium salt in grape juice, from which it is extracted and sold as cream-of-tartar, contains two hydroxyl and two carboxyl groups, COOH.CHOH.CHOH.COOH. *Citric acid*, to which oranges and lemons owe their sour taste, has one hydroxyl and three carboxyl groups.

Organic acids are neutralized by inorganic bases, forming salts,

$$CH_3.COOH + NaOH = CH_3COONa + H_2O.$$

A *dicarboxylic* acid, such as tartaric, takes two molecules of base for neutralization; a *tricarboxylic* acid takes three molecules of base.

$$COOH.CHOH.CHOH.COOH + 2\ NaOH = COONa.CHOH.CHOH.COONa + H_2O.$$

Here the hydroxyl, being alcoholic in its nature, is unacted upon by the base.

The salts are named by adding the termination *-ate* to the stem of the name of the acid. The salts of acetic acid are known as *acetates;* of . propionic acid, *propionates;* of tartaric acid, *tartrates,* and so on. The sodium salt of acetic acid is *sodium acetate,* the potassium salt *potassium acetate,* etc.

CHAPTER XV

Esters

Just as the organic acids react with bases to form salts they react with alcohols to form compounds known as *esters*.

$CH_3.COOH + C_2H_5OH = CH_3.COOC_2H_5 + H_2O.$

The esters correspond to salts, with an alkyl group taking the place of the metal of the salt. They are named like the salts with the name of the alkyl group of the alcohol written in place of that of a metal. So the ester given above is *ethyl acetate*. If methyl alcohol had been used instead of ethyl alcohol the resulting product would have the formula $CH_3.COOCH_3$, and would be called *methyl acetate*. If formic acid were substituted for acetic we would get methyl or ethyl *formate*, according to the alcohol employed. The esters are for the most part pleasant smelling liquids. The characteristic flavors of many fruits are due to the esters which they contain.

Under suitable conditions an ester will take up a molecule of water and reform the acid and alcohol from which it was derived.

$$CH_3.COOC_2H_5 + H_2O = CH_3COOH + C_2H_5OH.$$

This reaction is known as *hydrolysis*,[1] or *saponification*, and may be brought about by steam at high temperature, by alkalies, which combine with the acid as soon as produced to form a salt, or by the aid of small quantities of inorganic acids (H_2SO_4 or HCl), or by enzymes, which act as catalysts.

[1] See page 24.

More important than the simple esters, from a physiological standpoint, are the fats. *Fats* are esters of the polyhydric alcohol glycerine. As this alcohol has three hydroxyl groups it reacts with three molecules of acid. If these three molecules are all of the same acid we have a simple fat, whereas if they are different we have a *mixed* fat. If we represent any acid by R.COOH the reaction between this and glycerine to produce a fat will be written thus:

Glycerine	Acid	Fat
CH_2OH		$CH_2OOC.R.$
$CHOH$	$+ 3R.COOH =$	$CHOOC.R + 3H_2O.$
CH_2OH		$CH_2OOC.R.$

The most important fats are those produced by the interaction of glycerine and palmitic, stearic, and oleic acids. *Palmitic* and *stearic* acids have the formulæ $C_{15}H_{31}.COOH$ and $C_{17}H_{35}.COOH$ respectively, and belong to the same series as acetic acid. *Oleic*, $C_{17}H_{33}.COOH$, is an unsaturated acid with a double bond in its carbon chain. It has a lower melting point than the corresponding saturated acids, as is usually the case with unsaturated as compared with saturated compounds. For this reason fats in which the ester of oleic acid predominates are liquid, while those which are rich in palmitate or stearate are solid. Vegetable fats are usually rich in oleic ester, olive oil being about 95 per cent. of this fat, while animal fats contain a larger proportion of palmitate and stearate. The proportions differ somewhat in different species and even in different animals of the same species. The fat of cold-blooded animals is higher in oleate, and consequently lower-melting, than that of warm-blooded animals.

Pork fat contains about 67 per cent. oleic ester, human fat about the same amount, beef and mutton fat less.

When a fat is hydrolyzed, or saponified, by sodium or potassium hydroxide the products are glycerine and the sodium or potassium salts of the fatty acids.

$$\begin{array}{l} CH_2OOC.R \\ | \\ CHOOC.R \\ | \\ CH_2OOC.R \end{array} + 3NaOH = \begin{array}{l} CH_2OH. \\ | \\ CHOH \\ | \\ CH_2OH. \end{array} + 3R.COONa.$$

These salts are *soaps*. Properly speaking any salt of palmitic, stearic, or oleic acid is a soap, but the term is ordinarily restricted to the potassium and sodium salts, which are soluble in water, while all others (with the exception of the ammonium salts) are insoluble. Potassium salts are called *soft* soaps and sodium salts are *hard* soaps, although these terms are only relative, the hardness or softness depending also, to a certain extent at least, upon the amount of water left in the soap in the process of making. In homemade soaps and some varieties of commercial soap the glycerine is left in the soap, but more generally it is separated and saved, as it is a very valuable by-product, serving as the basis for many of our modern explosives.

Various theories have been put forward to account for the cleansing action of soap, of which only the one or two which appear to be most satisfactory will be mentioned here. According to the emulsification theory dirt is held to a dirty object by a thin layer of grease. Soap cleanses by emulsifying the grease, whereupon the dirt, no longer having anything to which to adhere, falls away. Doubtless this emulsification plays a part, possibly an important part, in the cleansing effect of soap. That it is not the whole explanation is proved by

experiments which showed that filter paper could be soiled by a suspension of lamp-black in water even when all grease had been carefully removed, while no soiling took place if soap were added to the water in which the lamp-black was suspended. To account for this it has been suggested that a combination of some kind occurs between the dirt and the object, and this is decomposed by the soap, which itself combines with the dirt, forming a soluble compound which is washed away. It is probable that no one theory is sufficient to explain the action of soap, but it may be that there is a combination of these two actions, possibly with some other as well.

When soap is added to "hard" water a curdy substance separates on the surface of the water, and no lather can be obtained until a large amount of soap has been added. This curd is a compound formed by the interaction of the soap with soluble calcium or magnesium salts which are present in the water.

$$2R.COONa + Ca(HCO_3)_2 = (R.COO)_2Ca + NaHCO_3.$$
$$2R.COONa + MgCl_2 = (R.COO)_2Mg + 2\ NaCl.$$

Not until all the calcium and magnesium salts have been removed from the solution can the soap exert its normal effect.

Since this action represents a considerable waste of soap, besides the formation of a disagreeable curd which is apt to stick to and stain clothes, utensils, etc., it is highly desirable that the salts to which it is due should be removed from the water before it is used for washing. This is what is known as "softening" water, and may be brought about in various ways according to the cause of the hardness.

Where the objectionable salt is calcium acid carbonate, as is frequently the case, it can be rendered insoluble, and therefore harmless, by boiling the water. Soluble

calcium acid carbonate is converted by heat into insoluble calcium carbonate, water and carbon dioxide being given off. (Compare the behavior of sodium acid carbonate, page 21).

$$Ca(HCO_3)_2 = CaCO_3 + H_2O + CO_2.$$

The scale which often forms inside a tea kettle is a deposit of calcium carbonate which has been precipitated from the water in this way.

Water which can be softened by boiling is said to have *temporary* hardness. If the hardness is due to calcium or magnesium chloride or sulphate it is said to have *permanent* hardness and boiling has no effect. These salts must be removed by precipitation with some other reagent. For household use the most satisfactory reagent for this purpose is sodium carbonate, (washing soda), which reacts to form the insoluble carbonates of calcium and magnesium along with soluble sodium salts which are harmless.

$$Na_2CO_3 + CaCl_2 = CaCO_3 + 2NaCl.$$

$$Na_2CO_3 + MgCl_2 = MgCO_3 + 2NaCl.$$

$$Na_2CO_3 + Ca(HCO_3)_2 = CaCO_3 + 2NaHCO_3.$$

Too large an excess of washing soda in the water should be avoided as it is injurious to fabrics, and indeed for most purposes. Its use in moderation, however, results in many cases in a very considerable saving. Different waters vary considerably in their hardness, the degree of hardness of any particular water being determined by the amount of standard soap solution which must be added to it before a lather will form. It is not uncommon to find water which uses up 20–40 ounces of soap per 100 gallons of water before any lather can be obtained, and if we allow the moderate estimate

of 10 gallons of water per person per day this would require from 50 to 100 pounds of soap per person per year for each person in the family. Paying at the present retail price of 20 cents a pound, a family of five would spend about $50.00 a year on the extra soap required to soften their water. Washing soda, on the other hand, is about three times as efficient as soap for this purpose, so that 17 pounds of washing soda at 3 cents a pound would effect as much as 50 pounds of soap at about one-twentieth the cost.

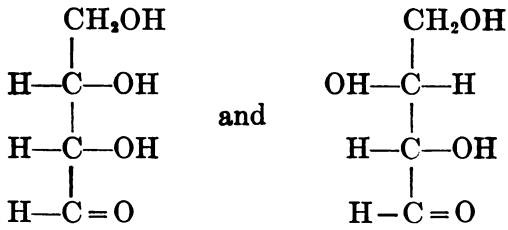
Where possible, it is well worth while having the water tested in order to find out the degree of hardness, and therefore the amount of soap and soda respectively required to soften it.

CHAPTER XVI

Carbohydrates

The *carbohydrates* are a very important class from a physiological standpoint, because so many of our important foodstuffs belong to this class. Carbohydrates contain an aldehydic or ketonic group attached to the last or the last but one carbon of a chain of carbon atoms each of which carries a hydroxyl group. *Sugars* are carbohydrates with a sweet taste. The carbohydrates are classified in groups according to the number of carbons which they contain. A *biose* has two carbon atoms, a *triose* three, a *tetrose* four, and so on, up to the *nonoses* with nine. Of these the *hexoses*, $C_6H_{12}O_6$, are by far the most important.

Among the carbohydrates there are a great many possibilities of isomerism, as will be seen from a consideration of the structural formulæ. If we consider the simplest carbohydrate, biose, there is only one possible arrangement of atoms which will answer the definition of a carbohydrate, $CH_2OH.CHO$, but we can have two trioses, an aldehydic or a ketonic derivative, $CH_2OH.CHOH.CHO$, and $CH_2OH.CO.CH_2OH$. When we come to the tetroses still further variation is possible. Not only may we have an aldehydic and a ketonic derivative, but if we write the formulæ of these in such a way as to indicate the relative positions of the various atoms in the molecule it will be seen that we must have different isomers according as the hydroxyls lie adjacent to each other or farther apart in the molecule.



The number of possible isomers increases very rapidly with increase in the number of carbon atoms in the chain, but comparatively few of these are of practical importance.

Carbohydrates with the general formula $C_nH_{2n}O_n$ are known as *monosaccharides*. This includes all the sugars from biose up to the nonoses. Three of these, all hexoses, are important food substances, *glucose* (also called *dextrose* or *grape sugar*), *fructose* (*lævulose* or *fruit sugar*), and *galactose*. Glucose and fructose differ only in being aldehydic and ketonic respectively, while galactose is aldehydic and differs from glucose in the relative arrangement of the H's and OH's in the carbon chain. Glucose is found in most fruits and some vegetables, and may be made from starch by a method to be discussed later. It is the component to which corn syrup owes its sweetness, having been formed from the starch of the corn. Fructose is also present in many fruits, but in smaller quantity than glucose for the most part. A considerable amount of this sugar is present in honey. Galactose is obtained from milk sugar.

Disaccharides are derived from two molecules of hexose with loss of one molecule of water.

$$2\ C_6H_{12}O_6 - H_2O = C_{12}H_{22}O_{11}.$$

The disaccharides can be hydrolyzed under the influence of acids or enzymes, when they take up water and sepa-

rate into two molecules of hexose,

$$C_{12}H_{22}O_{11} + H_2O = 2\ C_6H_{12}O_6$$

The reverse process, the building up of the disaccharides from the monosaccharides has so far never been accomplished in a practical way.

There are three disaccharides, *sucrose* (cane sugar), *maltose* (malt sugar), and *lactose* (milk sugar). One molecule of sucrose on hydrolysis gives one molecule of glucose and one molecule of fructose, the mixture of the two being known as invert sugar. From a molecule of maltose on hydrolysis we obtain two molecules of glucose, while lactose gives one molecule of glucose and one molecule of galactose.

A *polysaccharide* is formed from an indefinite number of molecules of hexose combined together with loss of water. Since it is impossible to ascertain the exact number of monosaccharide molecules united together it is customary to write the formulæ of the polysaccharides thus $(C_6H_{10}O_5)_x$, x representing an unknown number. On hydrolysis the polysaccharides are broken down in successive steps, passing through various stages of complexity until the monosaccharide is left. The most valuable of the polysaccharides is starch, the cheapest and most plentiful of our food-stuffs. When hydrolyzed by acids or by the amylolytic enzymes starch gives first a series of dextrins (polysaccharides of smaller molecular weight), then the disaccharide maltose, and finally glucose. It is in this way that the commercial glucose used as sweetening for cheap confectionery is made.

For the most part the reactions of the carbohydrates are those of the ketones and aldehydes. In the case of certain of the di- and polysaccharides however the aldehydic and ketonic groupings have been destroyed

in the union of the molecules, and the characteristic reactions of these groups can only be observed after the complex molecule has been hydrolyzed to the monosaccharides. This is true of the polysaccharides starch, dextrin, and glycogen, and also of the very important disaccharide sucrose (cane sugar).

The most important reaction of the sugars is their reducing action on the metallic oxides. A solution known as Fehling solution, which contains cupric ions, has long been used as a test for sugars in physiological fluids. Various modifications and improvements of this reagent have been devised but the general principal remains the same. The cupric ion imparts to the solution a blue color; on heating in the presence of a reducing sugar this color changes, passing through shades of green and yellow to red, owing to the reduction of cupric to cuprous ions, and finally a dark red precipitate of cuprous oxide, Cu_2O, settles out. This reaction, being very simple and accurate, is commonly used in testing for sugar in the urine. As it depends on the presence of the carbonyl group of the aldehydes or ketones, cane sugar and the polysaccharides will not react. For the most part any reducing sugar present in the urine is assumed to be dextrose, as this is the only sugar commonly present or of great physiological significance. When necessary specific tests can be used for further differentiation. By using a carefully prepared solution in which the quantities of reagent present are accurately known the estimation of reducing sugar by Fehling solution may be made quantitative instead of merely qualitative.

The carbohydrates are frequently classified according as they are or are not fermentable with yeast. *Fermentation* is a term used for decomposition brought about by the lower organisms through the enzymes

which they contain. The distinction between fermentation and other enzyme actions which involve decomposition of the substance acted on, rests on the fact that fermentation is a simple breaking up of the fermenting molecule itself into smaller molecules, whereas in other decompositions this breaking up is accompanied by combination with other atoms or groups to form the new molecules. This distinction will be clearer on comparing two typical reactions of each kind. The fermentation of dextrose to produce alcohol is represented by the following equation

$$C_6H_{12}O_6 = 2C_2H_5OH + 2CO_2,$$

while hydrolytic decomposition of a disaccharide to a monosaccharide is accompanied by the addition of a molecule of water.

$$C_{12}H_{22}O_{11} + H_2O = 2C_6H_{12}O_6.$$

Owing to the presence of fermentative bacteria in the intestines digestion is apt to be accompanied by more or less fermentation, the amount depending partly on the numbers of these bacteria present and partly upon the kind of food eaten. This fermentation leads to the production of irritating acids and gases and may cause extreme discomfort when allowed to become excessive.

The most important fermentation from a practical view point is the alcoholic fermentation of dextrose just referred to. To bring about this fermentation yeast is allowed to grow in a solution of the sugar. The yeast contains an enzyme *zymase*, which ferments the sugar, the energy set free in the reaction being utilized by the yeast plant in its growth. It used to be supposed that this fermentation was part of the vital process of the plant and therefore could not be carried on except through the agency of the living cell, but Buchner in

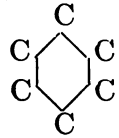
1897 succeeded in fermenting sugar by a yeast extract which had been so treated that it was impossible that a single living cell could have existed in it.

Zymase acts only on monosaccharides, but yeast contains another enzyme, invertase, which hydrolizes cane sugar to dextrose and fructose, thus preparing it for the action of the zymase, so that cane sugar solution makes a good medium for fermentation. Starch cannot be fermented by yeast, but the diastatic enzymes present in plant cells and in the digestive juices break it down into maltose, which is acted upon by maltase, another enzyme in the yeast cell, and hydrolized to dextrose, which is then fermented by the zymase. Starch can therefore be used as a source of alcohol by allowing these different enzymes to act upon it in turn. It is in this way that most of the alcohol of commerce is prepared, both for use as a beverage and for industrial purposes.

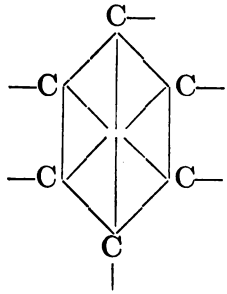
CHAPTER XVII

Aromatic Compounds

The *aromatic* compounds differ from the aliphatic in constitution and in general properties. Originally the term was applied to a small group of substances with a pleasant aromatic odor. Later many other substances were added on the ground of similarity in chemical properties rather than odor, until now the group is larger than that of aliphatic compounds and the distinction is based on structure only. The great host of *coal-tar products*, so-called because they are obtained from the tar left behind on the distillation of coal, belong for the most part to the aromatic series. Flavoring extracts, perfumes, dyes, medicines, and some of our most powerful explosives are among the many substances obtained directly or indirectly from this tar which is a mine of wealth to the manufacturing chemist.

The aromatic compounds have as their mother substance six carbon atoms arranged in a ring, or, as it is usually written for convenience, a hexagon, 

Each of these carbons has two valences used up by its neighbors on either side, a third holds a monovalent atom or group, and the fourth is supposed to be directed toward the centre of the ring, where the six meeting together are held in equilibrium by their mutual attrac-



tion, . The strength of this attraction is such that this carbon ring is perfectly stable, behaving in all respects like a saturated compound in spite of these apparently free valences.

The simple hydrocarbon, C_6H_6, with each carbon joined to one hydrogen atom, is called *benzene*, and the aromatic compounds may be looked upon as derived from this by replacement of one or more H's by other atoms or groups. For this reason the six-carbon ring is spoken of as the *benzene ring* or *nucleus*. The radicle C_6H_5, corresponding to the radicles CH_3, C_2H_5, etc. is called the *phenyl* radicle. The higher hydrocarbons are formed by substituting alkyl groups for hydrogen in benzene, as for instance *toluene*, $C_6H_5.CH_3$; *xylene*, $C_6H_4(CH_3)_2$; *ethylbenzene*, $C_6H_5.C_2H_5$. The alkyl groups thus introduced are spoken of as *side chains*. There may be one or more side chains in the ring, and these may be of any length, according to the size of the group introduced, and may be saturated or unsaturated. Styrene, $C_6H_5.CH{:}CH_2$, is an aromatic compound with an unsaturated side-chain.

It would be far beyond the scope of this book to deal with the aromatic compounds in detail. A few general characteristics will be pointed out, but for further information a larger work must be consulted.

Aromatic compounds exist which have the phenyl radicle combined with halogen to give halogen derivatives, with hydroxyl to give alcohols, with carbonyl to form

ketones and aldehydes, and with carboxyl to form acids, and these groups may be directly attached to the nucleus or may be separated from it by a longer or shorter chain of carbon atoms. While the properties of all these characteristic groups are somewhat modified by the benzene ring to which they are attached, they resemble the corresponding aliphatic compounds sufficiently in their general reactions so that no special attention need be given to them here. It may be taken as a general rule that side-chains and their substituent groups behave like aliphatic compounds, whereas the nucleus and any substituent introduced into the nucleus directly will show aromatic properties. This is particularly evident in the case of compounds containing hydroxyl, so much so that compounds with hydroxyl in the nucleus are classed together under the name of *phenols*, and only those with the hydroxyl in the side-chain of an aromatic nucleus are called *aromatic alcohols*. The phenols, while showing certain alcoholic characteristics, are so influenced by the presence of the benzene nucleus that they behave like weak acids, forming salts by replacement of the H of the OH with Na or K, *e.g.* C_6H_5ONa, *sodium phenolate*. On account of this acid characteristic the first member of the series, *phenol*, was supposed, when first discovered, to be an acid, and was named *carbolic acid*, a name which is still retained in general use although chemically incorrect. The phenols in general have antiseptic properties, and some of the higher members, as well as carbolic acid, are used as antiseptics, preservatives, and deodorants.

CHAPTER XVIII

Proteins and Vitamines

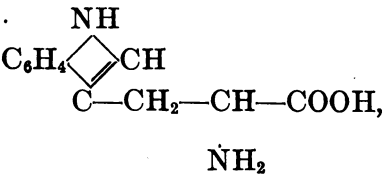
One of the most important, and one of the most complex, groups of substances with which the chemist has to deal is the group of proteins.

On heating any animal fluid or tissue extract an insoluble substance is obtained as a precipitate. If this precipitate be carefully dried and analysed it will be found to consist of one or more members of a well defined group of substances of similar chemical and physical properties which are classed together as *proteins*. The proteins all contain as essential constituents C, H, O, N, and S, are all built up on the same chemical principle, and therefore have a number of reactions in common. They have in addition certain specific characteristics which make it possible to subdivide them into smaller groups. Their molecules are so complex that it is impossible to assign any simple formulæ to them, and the only information we can obtain as to their structure comes from studying the root substances of which they are built up.

Proteins, like the polysaccharides, can be hydrolyzed by inorganic acids, alkalies, or suitable enzymes. Of these methods enzyme action is the most advantageous, since it goes on at ordinary temperatures, is less strenuous and therefore less apt to lead to undesirable decomposition products, and is easier to regulate in such a way as to permit of the isolation of intermediate products. Careful work has shown that hydrolysis takes place in steps, giving products of gradually decreasing complexity

until ultimately we obtain a mixture of simple compounds, all of one type, known as *amino acids*.

The amino acids are derived from the aliphatic acids such as acetic acid by the introduction of various substituent groups into the molecule. They are called amino acids from the fact that they all contain the *amino group*, $-NH_2$, attached to the carbon atom nearest to the carboxyl group. The simplest amino acid is *glycine*, $\overset{\overset{NH_2}{|}}{CH_2}.COOH$; others may contain longer carbon chains, such as *leucine*, $(CH_3)_2.CH.CH_2.\overset{\overset{NH_2}{|}}{C}H.COOH$; may have other substituents present, as *tyrosine*, $OH.C_6H_4.CH_2\overset{\overset{NH_2}{|}}{C}H.COOH$ with a phenolic group in the molecule, or tryptophane,



containing what is known as the *indole ring*; or may have more than one amino group present, as *arginine*, $HN{:}\overset{\overset{NH_2}{|}}{C}NH.CH_2.CH_2.CH_2.\overset{\overset{NH_2}{|}}{C}H.COOH$. One has sulphur present, *cystine*, $COOH.\overset{\overset{NH_2}{|}}{C}H.CH_2.S.S.CH_2.\overset{\overset{NH_2}{|}}{C}H.COOH$.

Many of these acids are combined together in varying proportions to form one protein molecule, the separate amino acids having somewhat the same relationship to the protein as the bricks and stones of which a building is made to the complete edifice. Just as an almost infinite number of different styles of architecture can be produced from a few simple types of building material so proteins are found in the most diverse shapes and forms. White of egg, lean meat, the curd of milk, nuts of all kinds, hair, and horn are all types of protein substances.

The importance of protein as food substance will be discussed later. Here it is sufficient to remark that the indispensable element which the proteins supply is the nitrogen atom present in the amino group, this being apparently the only form in which nitrogen is assimilable to any extent by the animal organisms. Plants have the power of synthesizing protein from the nitrogen of inorganic salts, carbon dioxide, and water, but animals lack this ability and are therefore dependent on what they can get from the plants, either directly or through the medium of other herbivorous animals.

Vitamines.—It has been found that a diet which contains sufficient fat, carbohydrate, protein, and mineral matter may still be insufficient for growth or even for the maintenance of the body in a healthy condition. Sailors on long voyages, deprived of fresh fruits and vegetables, the inhabitants of certain parts of the orient, whose diet consists almost entirely of polished rice, and babies fed exclusively on sterilized milk, are all subject to certain diseases which can be cured either by introducing other foods into the dietary or by the administration of small quantities of substances extracted from rice, bran, yeast, or orange and lemon juice. Diseases of this nature, of which scurvy and beri-beri are typical, are called deficiency diseases, and to the curative substances contained in fruits and vegetables was given the name *vitamines*. Recently this name has been extended to include certain growth-promoting substances of which some are found in fruits, vegetables, and the outer coats of cereals, and some in certain fats, especially the fat of egg yolk, butter, cod-liver oil, and beef suet. Both the chemical nature and the physiological effect of the vitamines are still somewhat obscure, but their importance in the diet is well recognized.

SECTION III

PHYSIOLOGICAL CHEMISTRY

CHAPTER XIX

Digestion

In order to maintain a living animal in health and activity it must be supplied with enough material to provide for the repair and rebuilding of the worn-out tissues, for the energy which it uses up in its various movements, and for growth in the case of the young. All this must come from the food, and the process of *digestion* consists in breaking down the various complex food substances into simpler soluble forms of which the organism can make use, and is followed by the selection and absorption of the useful portions, and the rejection and excretion of the useless material. The *assimilation* of the food is the process of building up these simple decomposition products of the food into the tissues of the body and the stores of reserve material which can be used to supply energy as needed. The whole series of processes through which the organism goes in the course of its vital activities, the breaking down of complex products into simpler ones and the building up of simple into more complex substances, is called *metabolism.* Metabolism includes not only digestion and assimilation but all the chemical reactions which take place in the body of a living animal.

The essential food for any animal must contain fats, carbohydrates, and proteins (these three being grouped

together as the organic food-stuffs), water, various mineral salts, and small quantities of organic acids which are beneficial if not absolutely essential, along with the growth promoting substances called vitamines. Most substances used as food are composed of a mixture of all these in varying proportions. It is the business of the dietician to determine what are the ideal proportions in order to obtain the best results and what food or combination of foods is best suited to the human system. Before any conclusions can be reached with regard to this it is necessary to follow the course of the food-stuffs through digestion and assimilation in order to determine the function of each and the factors which add to or detract from its efficiency.

In the mouth the food is crushed and torn apart by the teeth into small particles, not only in order to facilitate swallowing but also to expose as large a surface as possible to the digestive juices and to aid in solution (see page 28). It is also mixed with the *saliva*, the first of these digestive juices with which it comes in contact. The saliva is the mixed secretion of the glands of the mouth, of which the largest and most important are the *parotid*, which opens into the cheek just opposite the second molar, the *submaxillary*, which opens between the two halves of the lower jaw bone, and the *sublingual*, under the tongue, just behind the submaxillary. The secretion of these three differs somewhat, the parotid producing a thin watery fluid, while the lower glands give a thicker, slimy secretion which serves to lubricate the food particles so that they slip down the throat more easily.

The salivary glands secrete more or less continually, but they are stimulated to greater activity by various means, mechanical pressure (holding a stone in the mouth as athletes sometimes do to prevent their mouths and

throats from becoming dry), chemical stimulus (from the soluble substances in the food), and by the psychic influences, odor, sight, or even thought of appetising food: while on the other hand excitement, fear, anger, or embarrassment prevent the flow of saliva, so that the mouth and throat become dry and parched and the swallowing of food is rendered difficult.

Besides certain inorganic salts of which the exact function is unknown the saliva contains enzymes and a large amount of water. The water aids in the swallowing of food, helps to dissolve the soluble material, and when reabsorbed through the intestinal walls carries the dissolved substances with it into circulation through the blood stream, from which they can be taken by the various tissues as required. The most important enzyme present is the starch-hydrolysing enzyme *ptyalin*, or salivary diastase as it is sometimes called from its close resemblance to diastase, the starch-hydrolysing enzyme of the barley grain. There are also small amounts of maltase, which converts maltose into dextrose, and a proteolytic enzyme, but these are relatively unimportant.

The main function of the saliva appears to be the breaking down of the starch of the food, chiefly into maltose although a certain amount is further hydrolysed to dextrose. This change accounts for the fact that starchy food grows slightly sweeter as it is chewed. It is to be noted that while saliva acts readily on cooked starch its action on raw starch is very slow, a fact which shows the importance of thorough cooking of all starchy food so that it may be properly digested. Moreover thorough chewing in the mouth is of even greater importance as it is necessary that the food should be well saturated with saliva in order that the action may be as complete as possible. Long mastication ensures a copious flow of saliva as well as effecting a more complete

mixing with the food. It is a curious fact that, quite apart from the action on carbohydrates, all food digests faster in the stomach if mixed with saliva.

Although the action of the saliva in the mouth itself is only superficial it continues after the food has passed down into the stomach, during the long, gradual process of mixing with the stomach juice. When that mixing has been accomplished salivary digestion is stopped since the stomach juice is too strongly acid to allow the continued activity of the ptyalin.

The stomach is an elastic pocket, large at one end, small at the other, lying between the esophagus and the intestines and serving as a store-house where food can be retained and where it undergoes a certain amount of preliminary digestion, passing into the intestines in small amounts at a time for the completion of the process. The smaller end is called the *pyloric portion*, from the pylorus or thick ring of muscular tissue through which it passes into the small intestine, while the larger end, connecting with the esophagus, is distinguished as the *fundus*, or *cardiac portion*. The lining of the stomach is a soft mucous membrane, somewhat like the inside of the cheek to the touch, but lying in folds. All through these folds are small glands by which the *gastric juice* is produced and poured into the stomach.

The walls of the stomach being perfectly elastic, are able to contract and expand in proportion to the amount of food present. In the absence of food the walls collapse until they are in contact, but as the food passes down the esophagus the fundus becomes gradually distended, the food which enters first lying next the wall while the remainder is packed in the centre layer by layer. Properly speaking, therefore, the normal stomach is never empty, since there is never an unfilled space. There is a certain diseased condition, however,

in which the stomach remains permanently dilated in the absence of food, a condition which causes much distress to the individual.

The stomach glands, like the salivary glands, may be stimulated by various influences. Certain chemicals, particularly the soluble extractives of the food, cause the immediate production of juice. The extractives of meat are most effective, hence the value of a clear soup as an appetizer. More effective still however are the psychic influences, pleasurable anticipation causing a copious flow of juice while anger or excitement hinders or even prevents it. This was very well shown by the experiments of the physiologist, Pawlow, who operated on dogs in such a way as to produce a passage into the throat and the stomach through which food could enter or leave without the dog's knowledge. Through the stomach opening he could observe the progress of digestion under various conditions. He found that the flow of gastric juice began as soon as a plate of food was set before the dog and continued while he ate, whether the food passed into the stomach or dropped out through the esophageal opening, but stopped immediately if the dog was confronted with a cat. A somewhat similar observation was made on a child and it was found that when he was teased or distressed in any way the secretion of his stomach glands stopped. In another experiment two dogs were given pieces of meat of equal size, but one dog was allowed to eat his meat in the ordinary way, while in the other case it was inserted through the stomach opening without the dog's knowledge. At the end of an hour or two the meat in the stomach of the first dog was found to be much better digested than that in the stomach of the second where no psychic influence came into play.

Where food is unappetising it is not entirely undigested

because the chemical influences will still have some effect, but digestion is much slower owing to the smaller supply of juice from the glands. It is highly desirable therefore that food should be not only nutritious, but also attractively served, eaten slowly, so as to ensure thorough mastication, and, as far as can be secured, amid pleasant surroundings. As anger and emotion inhibit the action of the glands all disagreeable or exciting topics should be avoided at meal time. In the case of children it has been found that even vigorous play just before a meal has a bad effect on the activity of the stomach glands and should therefore be prevented. While all these have been long recognized more or less as good general rules of conduct it was only by actual experiment that they could be shown to be based on sound physiological reasons.

Pure gastric juice is a colorless liquid looking much like water, with a slightly sour taste due to a little hydrochloric acid which is always present to the extent of about 0.2–0.4 per cent. in the gastric juice of a normal individual. In much larger amounts than this it interferes with digestion, and the condition is said to be one of *hyperacidity*. *Hypoacidity* on the other hand is the condition where there is too little acid in the gastric juice, and is equally injurious. The principal enzyme present is *pepsin*, which acts upon proteins, bringing about the first step in their decomposition and producing certain slightly less complex, soluble products known as *peptones* and *proteoses*. Pepsin as secreted by the glands is in the inactive zymogen state (see page 51), and only becomes active in the presence of hydrochloric acid, so that the activating of this enzyme is one, although probably not the only, function of the acid in the stomach. Besides pepsin we find two other enzymes present, *lipase*, or fat splitting enzyme, and *rennin*, an enzyme

which coagulates milk into a smooth curd. Thanks to the lipase present gastric juice has considerable fat-digesting power, provided the fat is sufficiently finely divided. From 50–80 per cent. of the fats in milk and egg yolk, where they are present as an emulsion, are hydrolyzed to fatty acids and glycerine in the stomach, while 98–99.5 per cent. of fats such as those of meats and butter, which are not emulsified, pass unchanged into the intestines.

Various commercial preparations of rennin (or rennet, as it is sometimes called) extracted from the stomachs of calves are on the market and are used for coagulating slightly sweetened and flavored milk. The curd thus produced is called junket and is a palatable and nutritious food. It seems probable that this coagulating or clotting of the milk is due to a partial hydrolysis of the *casein* (the protein constituent of the milk) into a simpler and less soluble form to which the name *paracasein* has been given. It is, then, a preliminary digestion analogous to that brought about by pepsin when proteins are hydrolysed to peptones and proteoses.

In the normal stomach the contents of the pyloric region are continually being churned up by the periodic contraction of the stomach wall with a kind of wave motion, starting at the slight constriction between fundus and pylorus and running towards the pylorus. This motion at first very slight, becomes more vigorous as digestion proceeds and the contents of the pyloric region, under the combined effect of churning and enzyme action and the solvent power of the juice, becomes transformed into a semi-fluid mass called *chyme*. Finally, when the glands have secreted enough juice to give a distinct acid reaction to the chyme, the pylorus relaxes and the next wave of contraction forces a portion of the chyme through into the intestine. The pylorus then

closes, a portion of the food mass in the fundus is pressed toward the pylorus and digestion continues. The length of time required to completely empty the stomach varies from one to seven hours according to the quantity and character of the food eaten. Protein remains longer in the stomach than carbohydrate, fat longer than protein, and mixtures of fat and protein longer than either alone.

In considering the digestibility of a food a distinction should be made between the *ease* of digestion, that is the rapidity with which the food leaves the stomach, and completeness of digestion. In general the more fluid the food the more quickly will it pass through the stomach, but it does not necessarily follow that it is more digestible, in the sense of providing more nutriment, than a food which remains in the stomach longer. Fluid white of egg is more quickly digested, and hence is usually considered more digestible, but more nutriment is obtained from hard-boiled egg white. In cases of weak digestion it is often desirable to sacrifice completeness for ease of digestion, although where the digestion is good it is more economical to give those foods which are digested as completely as possible.

Some, but probably a very small proportion, of the food is directly absorbed into the system through the stomach walls. Water is so absorbed, as are also certain drugs and poisons, as is shown by the rapidity with which they act on the system. Alcohol is directly absorbed and permeates the whole system with great rapidity, showing its effects within a very few minutes after being swallowed. The bulk of the food, however, reaches the intestines, where it undergoes further digestion and is finally absorbed through the intestinal wall.

The small intestine is a tube about thirty feet long coiled in the interior of the body cavity. The *intestinal*

juice is the combined secretion of the *intestinal glands*, a large number of small glands which are contained in the intestinal wall, the *pancreas*, a very large gland opening into the intestine four or five inches below the pylorus, and the *liver*.

The pancreas, like the other digestive glands, produces its secretion only in response to definite stimulus. When the acid chyme arrives in the intestine it causes the production in the intestinal wall of a substance called *secretin*, which acts upon the pancreas in such a way as to bring about secretion of the pancreatic juice. This juice and that of the smaller intestinal glands is strongly alkaline due to the presence of carbonates. This alkali serves to neutralize the acid of the chyme, and also reacts with the fatty acids produced by the hydrolysis of the fats, forming soaps (see page 91) which help to emulsify the still undecomposed fats and render them more susceptible to the attack of the lipolytic enzyme.

There are several enzymes present in the pancreatic juice, of which the most important are *steapsin*, which hydrolyses fats to glycerine and fatty acids, *amylopsin*, which, like ptyalin, converts starch into maltose and therefore supplements the action of the saliva, and *trypsin* which acts on proteins, converting them into amino acids. Besides these there are also *lactase*, which hydrolyses lactose into dextrose and galactose, *maltase*, which hydrolyses maltose into dextrose and at least occasionally *invertin* which hydrolyses cane sugar into dextrose and fructose.

The secretion of the small intestinal glands is very like that of the pancreas in some respects. It contains invertin, maltase, and lactase, and in addition an *enzyme*, *erepsin*, which acts on the peptones and proteoses produced by the pepsin in the stomach, completing their digestion into amino acids; and still another enzyme *enterokinase*,

which serves to activate the trypsin, possibly in somewhat the same way in which pepsin is activated by the hydrochloric acid of the gastric juice. If collected alone, as directly secreted by the gland, pancreatic juice has little or no action on proteins, but when mixed with the juice from the intestinal glands it acts vigorously. On the other hand, the juice from the intestinal glands alone has very weak digestive action, but seems to serve mainly to strengthen the action of the pancreatic juice.

The liver differs from the other digestive glands in producing a more or less continual secretion which is stored up in the gall-bladder and is poured from this into the intestine during digestion. This secretion, known as the *bile*, is a thick, viscous fluid, golden or greenish yellow in color. When for any reason the bile duct becomes stopped up the disorder known as *jaundice* follows, and the coloring matter of the bile becomes distributed through the body, giving a yellowish tinge to the skin. *Gall-stones* are deposits of solid matter which for some reason have separated out from the bile in the gall-bladder, where they sometimes cause great suffering. While the bile in itself has little or no digestive action, it has a remarkable effect in assisting the action of the other secretions, especially in facilitating the digestion and absorption of fats. Not only has it great emulsifying powers, being even more effective than the soaps in this respect, but it seems to be able to enter into some peculiar chemical or physical combination with the fatty acids or their salts, in which form they readily pass through the intestinal walls into the lymph vessels, from which they are carried through the system. In the absence of bile very little, if any, absorption of fat takes place.

After the chyme has passed out of the stomach it lies for some time in the upper loop of the intestine, while the

intestinal juices are poured over it and the work of digestion continues. After a time a churning motion of the intestine begins, somewhat like the contractions of the stomach, and the food is divided into small portions and vigorously churned. This motion, besides mixing the food with the digestive juice forces it into closer contact with the intestinal wall and hence increases the absorption from the intestine. This motion continues for some time and then stops, to be succeeded by a slower wave motion which travels down the intestine pushing the food down to the lower end. This motion is called *peristalsis*, and is of great importance. Where the tone of the intestinal wall is poor and peristalsis is weak constipation results. Certain drugs, such as calomel have the power of restoring the lost tone and stimulating peristalsis and this appears to be a function of the bile also. Mechanical stimulus seems likewise effective, hence the value of a certain amount of coarse, undigestible material such as the cellulose of vegetables or the bran of cereals in the diet. These substances are insoluble and impervious to the digestive juice of human beings, but the pressure which they exert on the walls as they pass through the intestinal tract is stimulating, peristalsis is increased, and hence a tendency toward constipation may be overcome.

As the food moves down the digestive tract it meets no new secretions, but the juice with which it is already saturated continues its action and absorption takes place gradually. From the small intestine it passes into the large intestine, or colon, where it may remain a day or even longer before being excreted from the body. Absorption continues in the upper part of the colon, but decreases very considerably until in the lower part there is none at all. There are no secreting glands in the colon and hence no digestive juice except such as the food

brings with it, but there are large numbers of bacteria which act on the food producing various more or less injurious decomposition products.

There are three main types of bacteria which inhabit the human digestive tract, the *putrefactive bacteria*, the *acid forming bacteria*, and bacteria of the class known as the *B. Coli type*. Of these the putrefactive bacteria seem to be the only forms which are very actively injurious. They produce poisonous decomposition products which are absorbed and penetrate the system causing many more or less marked symptoms ranging from mental depression, drowsiness, and irritability, generally accompanied by physical discomfort, headache, outbreak of pimples, cold sores, boils, development of colds or aggrevation of mild chronic ailments, up to the acute condition recognized as ptomaine poisoning. The longer the food residue remains in the colon the greater the opportunity for its decomposition by the putrefactive bacteria, hence one of the great evils of constipation. Much interest was aroused a few years ago by the theory of a Russian physiologist, Metchnikoff, who regarded old age with its failing powers and decreased resistance as the result of gradual poisoning of the whole system by these products of putrefaction and suggested that the secret of long life and health lay in eliminating the injurious organisms. As it is of course impossible to sterilize the digestive tract of an animal or even to render it less favorable for the bacterial growth the only possible method of disposing of the putrefactive bacteria is by taking advantage of the fact that when two races of bacteria are growing in a limited space the stronger and more vigorous will gradually crowd out the weaker, until finally the latter disappears altgoether. Metchnikoff therefore sought for some type of bacteria which would be antagonistic to the putrefactive bacteria

and which could be cultivated in the intestines without themselves producing injurious effects, and he believed he had found this in the *lactic acid bacteria* present in sour milk. Accordingly he recommended the drinking of sour milk or buttermilk containing large numbers of lactic acid bacteria, especially those of a certain strain, *B. Bulgaricus*, which could be developed by inocculating sterilized milk with the culture and keeping it for a few hours in a warm place. The buttermilk cure became the fad of the moment, but died out after a time as all dietetic fads are sure to do. Whether buttermilk is or is not a preventive of old age it is unquestionably a wholesome and nutritious drink and its extensive use might well be encouraged, but perhaps an easier and surer method of reducing intestinal putrefaction is by careful diet, especially by limiting the amount of meat eaten and by very thorough mastication so that the least possible amount of undigested protein reaches the region of bacterial activity. An increase in the proportion of carbohydrate food also tends to decrease putrefaction, partly because the bacteria will not attack the protein in presence of sufficient excess of carbohydrate, and partly perhaps because the carbohydrates in their fermentation produce acid which is unfavorable to the growth of many forms of bacteria.

While bacteria flourish in the colon and the lower part of the small intestine they are not found in the upper part of the intestine nor in the stomach; in fact bacteria introduced through the mouth are usually destroyed by the acid gastric juice, which has a strong antiseptic effect.

Summary.—The three digestive juices are the saliva of the mouth, the gastric juice of the stomach, and the intestinal juice, which includes the secretion of the small intestinal glands, the pancreatic juice, and the bile. The combined effect of all these agents is to decompose

the complex and insoluble food materials into simple molecules which are soluble and which can be absorbed and either decomposed further to supply energy or built up into body substance by the various organs. Carbohydrates, except the indigestible cellulose, are hydrolysed to monosaccharides, the action taking place chiefly in the stomach but being completed in the intestines. Fats, if very finely divided, are hydrolysed in the stomach, but for the most part they pass into the intestine where they undergo decomposition into fatty acids and glycerine. Protein undergoes a partial decomposition in the stomach and the products of this action, as well as any protein which has passed through the stomach unchanged, are hydrolysed to the amino acids in the intestines. The mixture of food and digestion products is pushed slowly down through the intestines while absorption takes place gradually. In the lower end of the small intestine and the colon the mass is subjected to the action of various bacteria, of which at least one race, the putrefactive bacteria, produce injurious substances which are absorbed and spread through the system.

CHAPTER XX

Assimilation

Fats.—In the course of their passage through the intestinal wall the fatty acids and glycerine become recombined into fats again, although the character of the fat is usually altered somewhat, fats of all kinds taken in the food, vegetable fats and oil, butter, pork fat, beef fat, being all alike converted into the peculiar body fat of the animal concerned. This fat is taken up from the intestinal wall by the lymph vessels and poured directly into the blood, by which it is transported to the tissues, where the greater part of it is deposited as reserve material, though a certain amount may be combined into cell substance or burned directly if needed for energy. Fat is a very convenient form in which to lay up reserve food, since it can be deposited fairly uniformly over the body, causing no inconvenience, except when present in excessive amounts, and can be readily reconverted into a soluble or diffusible form for rapid transport to any part of the body where it may be needed at any time. It is the most economical fuel available to the body, giving about twice as much heat per unit weight as either carbohydrate or protein. Moreover it is valuable as a covering and protection for the nerve endings and for certain other delicate organs. In order that the fat reserve may be very quickly available for the organism fatty tissue is permeated with a network of minute blood vessels, and if there is too great an extent of such tissue the effort of pumping blood through all these small vessels may put an undue strain on the heart.

Where there is a tendency toward excessive storage of fat in the tissues it should be guarded against by careful dieting, both as to kind and amount of food taken, and by excercise, which necessitates the burning of more fat to supply the energy. Where ordinary care in this respect is insufficient to counteract this tendency the advice of a competent physician should be sought, as incalculable harm may be done by some of the drastic methods of "cure" by starvation or drugs used at random.

Proteins.—The amino acids produced from the proteins by digestion are absorbed into the blood stream and carried to the tissues. From these amino acids the body protein is built up, the cells being the builders as well as the firemen who prepare the fuel and keep up the fires which supply the necessary energy for the vital activities. They must also aid in the destruction of unnecessary material which would clog up the system and have an injurious effect if allowed to accumulate. In the young growing organism they must construct new tissue to add to the body structure, but in the adult it is only necessary to repair the waste and maintain the efficiency of the living machine. Their task is complicated by the fact that they have only second-hand materials with which to work and that the supply of those is too often regulated by the uncontrolled appetite of the body they serve, with little or no reference to the actual needs of the system. In such a case the cells must make what shift they can, destroying or excreting the surplus thrust upon them, sifting out the necessary amino acids from the rubbish, and if necessary tearing down and burning the body substance itself in order to maintain the life-giving fires.

Not all the amino acids are of equal importance; some can be made in the body from other materials,

others are required only in small amounts, while still others are essential for growth or maintenance of life. An animal fed on maize as its sole protein food never attains more than two thirds of its normal growth, however abudant its food may be in all other respects, but if a small amount of milk be added to the diet normal growth is induced. Similar results are obtained with wheat and oats, but in the latter case gelatine proved a better supplmentary food than milk. The explanation is believed to be that while these proteins may contain all the necessary amino acids certain ones are present in very small quantity, while others are in excess. If some other protein can be found which is rich in the acids which the first lacks and poor in those which the first contains in abundance the combination of the two in reasonable amounts will be more satisfactory as a food than either alone. When our foods have been more thoroughly studied from this point of view we will doubtless be able to arrange various combinations of proteins which contain the correct proportions of all the amino acids, but in our present state of ignorance it seems to be the safest plan to include in the diet a variety of proteins in rather small amounts, rather than a large amount of any one, in order to obtain a good average of amino acids.

Lack of the fundamental amino acids is of course fatal for the building cells, since no other food material contains nitrogen and hence nothing else can be substituted for protein. On the other hand too much protein causes a great deal of unnecessary work in the disposal of it. Unlike carbohydrate and fat protein cannot be stored up to any extent as reserve material in the body. During active growth and occasionally after a long wasting illness or a hemorrhage which has very greatly reduced the amount of protein in the body, a considerable amount

may be retained in the tissues, but ordinarily no matter how much may be taken in it is practically all got rid of in the course of twenty-four hours. What becomes of this protein which passes through the system? The amino acids which are not needed in that form first undergo *deaminization* in the tissues, that is, the nitrogen-containing part of the molecule is split off from the rest and excreted promptly in the urine, while the remainder may be converted into carbohydrate or fat and either stored in the body or burned to carbon dioxide and water. As much as 58 per cent. of the total protein may be converted into carbohydrate, a fact which is of importance in diabetic feeding (see page 126). Protein may thus take the place of carbohydrate or fat as a source of energy, apart from its function of tissue building, but it must be remembered that its use for that purpose is wasteful both of material and cell activity, since so large a proportion of the protein molecule remains combined with the nitrogen and is excreted without giving up its energy to the system. Moreover there is reason to believe that the unnecessary strain on the organs engaged in deaminizing and excreting large quantities of protein is injurious.

Carbohydrates.—The monosaccharides pass into the blood on absorption and are carried directly to the liver where they are built up into a new carbohydrate complex, glycogen, which is stored up in the liver and decomposed into dextrose as required. The muscles also contain glycogen, and during exercise this glycogen is hydrolyzed to dextrose which is then further decomposed to carbon dioxide and water, setting free chemical energy which the muscles convert into mechanical energy. The supply of glycogen in the muscle is thus depleted, but the muscle supplies the deficiency by taking dextrose from the blood (which always contains about

1 per cent.) and building this up into more glycogen. As the blood becomes poorer in dextrose however, more of the glycogen of the liver is decomposed into dextrose which passes into the blood, only to be taken up by the muscles again. There is therefore an automatic regulation by which the supply of carbohydrate in the blood and muscles is kept constant by the liver. This organ only contains as a rule about one-third of a pound of carbohydrate, and about as much more is distributed through the muscles, along with a very small amount in the blood. The total carbohydrate content of a body at rest is therefore something less than one pound, and this amount may be greatly reduced by exercise or exposure to cold. The amount taken in the food must maintain this level in the body and supply any excess that may be called for by the day's activity. If more than this is taken in, under suitable conditions, it is transformed into fat and stored in that form; hence the well-known fattening power of carbohydrate food. On the other hand, if such a large excess of carbohydrate is taken that the system can neither use it as carbohydrate nor transform it sufficiently quickly into fat it is excreted through the kidneys and appears as dextrose in the urine, where it may readily be detected by the Fehling test (see page 98). This phenomenon is frequently observed soon after a large amount of carbohydrate has been eaten and is then called *alimentary glycosuria.* It is especially likely to occur if the carbohydrate of the food is present in the form of sugar, since starch requires a much longer time for its hydrolysis and is therefore absorbed more gradually. Dextrose is even more rapidly assimilated than are the disaccharides, and for this reason and because it is less sweet than cane sugar there is a little more danger of constant over-eating of this form of carbohydrate if it is used in unrestricted amounts.

This is the only legitimate objection that could be made to glucose as a food and is the chief reason for legislating against its indiscriminate use by confectioners and bakers.

While the above mentioned alimentary glycosuria is merely temporary and is of no consequence except inasmuch as it is a waste of food material, a chronic glycosuria or elimination of sugar is one of the symptoms of the disease known as *diabetes*. Accompanying symptoms are increase in the sugar content of the blood and appearance of acetone derivatives, chiefly of an acid character, in the urine, a condition which is known as *acidosis* and is believed to be due to the imperfect combustion of fats and proteins. It should be noted, however, that very fat people without any diabetic tendency are prone to acidosis, due to the excessive amount of fat which they decompose, while on the other hand very thin people show the same tendency on account of excessive decomposition of protein in their bodies.

The cause of diabetes is still undetermined, but enough has been learned to point the way to various forms of treatment which have been used with considerable success. Administration of sodium bicarbonate is found to be effective in counteracting the tendency to acidosis, while the excretion of sugar can be largely controlled by careful diet and exercise. In the early history of the disease it was believed sufficient to restrict the carbohydrate food and allow the patient to get all his nourishment from fats and proteins, but this treatment was only better than none at all. That it could not be really satisfactory will be evident when we consider that about 50 per cent. of the protein consumed may be converted into glucose and the remainder may go to increasing the amount of acetone bodies present, while fats, although they do not increase the glucose excretion, do increase

acidosis. There must therefore be a general reduction in the total food until the urine is sugar free, when the diet may be slowly increased. In the Allen treatment which has been used with notable success in this country the patient is first made to undergo a prolonged fast until the excretion of sugar stops completely. This rests the weakened functions and is found to be very beneficial even in cases where the patient is emaciated at the start. During this fasting period water is given freely, along with a little clear tea or coffee if desired. Some physicians prefer to give a little food after two or three days and then recommence the fast, others prescribe one unbroken fast until the desired result is obtained. The diet is then very carefully regulated, the food being increased very slowly, with frequent tests to show whether any sugar is being excreted. At the reappearance of a trace of sugar the fast is repeated until it disappears again and the feeding is then begun at a lower level than the point at which it was stopped and increased still more gradually. Under these conditions a *tolerance* is established first for carbohydrate, then for protein, and lastly for fat. After this point has been reached reasonable care and self control on the part of the patient is usually sufficient to keep the disease latent, but over feeding will not only cause the reappearance of sugar but also acidosis and the condition may be even worse than before treatment.

For the normal person sugar, owing to the ease and rapidity with which it can be digested and assimilated, is a prompt and efficient stimulant, especially in case of muscular fatigue. The reviving effect of a cup of hot chocolate on a cold day is partly due to the warming up of the body by the hot liquid and partly to the rapidly absorbed carbohydrate which replaces the stores depleted by cold and exertion.

CHAPTER XXI

The Energy of the Body

Since any form of energy can be converted into any other form and since if the transformation be properly carried out there will be no loss of energy, it follows that a given amount of mechanical energy is equivalent to a given amount of heat energy and vice versa. So if we know how much heat a piece of coal will give out we can calculate how much water it will convert into steam and what load that steam will lift; in other words, we can express the heat value of the coal in terms of mechanical energy. On the other hand, when a hammer falls on an anvil it produces a certain amount of heat, the amount being proportional to the weight of the hammer and the height from which it falls, so that the mechanical energy of the blow can be expressed in terms of heat energy.

In order to measure anything it is necessary to have some unit of measurement. There are various units of weight in common use, the ounce, the gram, the grain, etc. We may measure length in inches, or feet, or yards. To express volume we have the pint, the liter, and so on. Similarly we measure amounts of heat in *Calories*. The Calorie is the amount of heat which will heat one pound of water just about 4°F.

Since the main function of the food is to provide energy for the organism it becomes of importance to be able to calculate the energy value of any food, and this can be done by ascertaining how much heat it will give out when it is decomposed in the body. Since the sum total

of the changes undergone by the food in the body results in its oxidation the problem resolves itself into the determination of how much heat a given amount of protein will give out when it is oxidized, or burned. This can be done by making use of an appliance called a *calorimeter* which consists of a steel tube in which a weighed amount of food is placed and which is then filled with oxygen and sealed up. This is immersed in an outer vessel filled with water and so arranged that no heat can get into it except from the inner tube, or out from it to the outside air. The food in the inner tube is then ignited by means of an electric spark and allowed to burn. The heat produced warms the water in the outside vessel and by weighing the water and noting the number of degrees of temperature change the number of Calories given off by the food in burning can be determined. Moreover, it has been found possible to devise an enlarged calorimeter on somewhat the same plan, in which a human being can be placed and where the heat given off from the body can be measured, as well as the carbon dioxide and water eliminated from the lungs and skin and the oxygen used up in metabolism. By means of this appliance it has been found that the heat given off in the oxidation of carbohydrate and fat is exactly the same whether they are burned in the body or outside it. The heat given off by a burning substance is called its *heat of combustion.* This heat is found to be about 113 Calories per ounce for carbohydrates and proteins and about 255 Calories per ounce for fats. Proteins give off a little less heat in the body than when burned outside it, owing to the excretion of incompletely oxidized products, but when a correction is made for these the agreement is truly remarkable, considering the difficulty of making accurate observations of this sort.

With the help of the calorimeter it is possible to find

out how much energy is being used up in various kinds of exercise, and even, by calculating the relative amounts of oxygen used up and of carbon dioxide and nitrogen compounds excreted, to determine accurately what kind of material is being oxidized to produce this energy. Many interesting and important facts about human metabolism have been learned. It has been proved that the body can obtain its energy from either carbohydrate, fat, or protein, and therefore these substances can replace each other in the diet to an extent proportional only to their energy value; that is, since an ounce of carbohydrate gives a little less than half as much heat as an ounce of fat, half an ounce of the latter will more than replace one ounce of carbohydrate. This is known as the *Law of Isodynamics*, and is of great importance to the dietician since it means that a diet may be widely varied according to taste or convenience, provided the energy, or *fuel value*, of the food remains unchanged. This law must, however, be used with certain restrictions. Neither fat nor carbohydrate, could be entirely substituted for protein, since neither of these contains the essential nitrogen which the tissues require. If fat and carbohydrate are lacking, the necessary energy can be obtained from protein alone, and on the other hand, if they are present in abundance much less protein is broken down, as is shown by the decreased amount of nitrogen compounds in the urine. Moreover, as has been already pointed out, two kinds of protein are not necessarily of equal value even though their fuel value is the same, since one may have a better proportion than the other of the amino acids which the tissues require.

CHAPTER XXII

The Blood

Just as the nerves may be regarded as the telephone or telegraph system by which messages are sent from one part of the body to another, so the blood corresponds to the express service by which materials of all kinds can be transported from place to place. Food materials must be carried from the digestive tract and oxygen from the lungs. Various organs produce secretions which stimulate or modify the action of other organs, and these secretions are taken by the blood to the point of application. Waste materials must be carried from the tissues to the excreting organs. The great convenience of a rapidly circulating medium like the blood which can dissolve some substances and absorb or combine with others to form compounds which are easily separated again is obvious.

The blood consists of a liquid portion, the *plasma*, in which float large numbers of living cells, the *corpuscles*. The plasma is a clear yellowish fluid of which about nine-tenths is water, and the remainder chiefly protein, with about 1 per cent. of inorganic salts and small quantities of other substances such as dextrose, fatty substances, nitrogen compounds, etc. Blood freshly drawn is perfectly fluid, but on standing it clots or coagulates in the course of from 2 to 10 minutes into a dark red jelly-like solid, from which a little yellowish liquid, *the serum*, separates, very like the plasma in appearance, but with lower protein content. The clot is found to consist of a very fine net-work of *fibrin*, a protein substance

separated from the plasma, in which the corpuscles are emmeshed and held. If instead of allowing the blood to stand quietly it is vigorously stirred or beaten the fibrin separates, but not in such a fine net, so that the corpuscles instead of being held in a clot remain suspended in the serum. This suspension of corpuscles in plasma from which the fibrin has been removed is called *defibrinated blood.* The clotting of blood is a property of greatest importance, since the formation of a clot stops bleeding and so prevents fatal hemorrhage. It seems to be of the nature of a crystallization of the fibrin, but authorities are undecided as to its cause. The presence of calcium salts is essential for clotting and the action is greatly hastened by contact with the wounded tissue or with some foreign substance which seems to act as the foundation on which the deposition of the fibrin can begin. The use of cobwebs to stop bleeding, as was done in olden days, depended on this principle, but was of course very dangerous owing to the likelihood of infecting the wound. Anæsthetics render the blood abnormally easy to clot, and hemorrhages have the same effect so that clotting in the veins of the legs, particularly in the right leg, is not unusual, for instance after parturition or an appendicitis operation. The condition known as "milk leg" is the result of such a clot becoming fixed and disturbing the circulation of the leg.

Hemophilia is a derangement of the blood which prevents clotting, so that people have been known to bleed to death from so small a wound as that made by pulling a tooth. It is inheritable, and, curiously, is usually transmitted by the females but appears in the male members of the family.

The *erythrocytes,* or red corpuscles of the blood, are small gelatinous discs, slightly hollowed out on the surface, with a diameter of about one-thirty-two hundredth

of an inch, and a thickness of about one-fifth as much. The blood of the adult male normally contains from 5–6,000,000 per c.mm., that of females from 4–4,500,000, but in anemia the number may be reduced to less than half of this. On the other hand they are increased by massage, by hot or cold baths, by the administration of certain drugs, and in diseases such as cholera, diarrhea, and dysentery.

The most important constituent of the red corpuscles is the hemoglobin, or red coloring matter, which has the power of combining with oxygen to form a very unstable compound which is called oxy-hemoglobin and which very readily gives up its oxygen to the tissues. Oxy-hemoglobin is bright red in color, while hemoglobin has a darker, purplish shade, hence the difference in color between blood flowing from the arteries, which convey the blood from the lungs to the tissues, and that from the veins, which carry it from the tissues to the lungs again. Hemoglobin is a very complex protein substance containing iron. With an insufficient amount of iron in the food formation of hemoglobin decreases. Addition of iron salts greatly influences the production of red corpuscles, and especially the amount of hemoglobin they contain, but their method of action is obscure. With regard to the oxygen carrying power of hemoglobin little can be added to what has already been said, that hemoglobin has the power of uniting with oxygen and giving it off again to the tissues (see page 17). That the oxygen is actually chemically combined and not merely dissolved in the blood is proved by the large amount of oxygen present in arterial blood as compared with the amount which will dissolve in the blood serum. Hemoglobin can combine with other gases also, notably carbon monoxide. When blood has been exposed to a mixture of this gas with oxygen or air the hemoglobin combines

with the carbon monoxide rather than with the oxygen, thus preventing the usual supply of oxygen from reaching the tissues. 50 per cent. saturation with carbon monoxide endangers the life of an animal, and with human beings this can occur in the presence of as little as 0.05 per cent. of carbon monoxide in the air. The fatal effect of prolonged exposure to coal gas is largely, if not entirely, due to the carbon monoxide present.

Owing to the vast number of corpuscles and their peculiar shape they expose a very large surface to the action of the oxygen. It is calculated that under normal conditions the combined area of the red corpuscles in an adult man would be about 3520 sq. yds. representing between one and two pounds of hemoglobin. The advantage of having this surface distributed among a great many smaller particles instead of among a smaller number of larger particles is partly because of the ease of transportation through the small passages and partly because of the great rapidity with which they can be loaded and unloaded with oxygen.

If the corpuscles are dissolved or injured so that the hemoglobin escapes into the plasma the blood is said to be *laked*, and the destruction of the corpuscles in this manner is called *hemolysis*. Hemolysis may be brought about by immersing the corpuscles in water or in a hypotonic salt solution (see page 26), or by the action of chemical substances, including the products of bacterial life and the poison of snake bites, spider and bee stings, etc.

The white corpuscles or *leucocytes* are somewhat larger than the red corpuscles and are present in the proportion of about one of the former to 350–500 of the latter. They behave like independent organisms and are capable of spontaneous movement, so that they can enter and leave the blood and penetrate the tissues at will. Their

most remarkable property is the power of absorbing solid particles of tissue, food substances, bacteria, etc. Since they assemble in large numbers around the digestive tract during the process of digestion it has been suggested that they may help to carry the products of digestion through the system, but their main function is believed to be the removal of old, worn-out or diseased tissues and cells, and to digest or destroy injurious organisms which find their way into the tissues. The power of the leucocytes to destroy bacteria in this way seems to be connected with the presence in the blood of certain substances called *opsonins*. The higher the *opsonic index*, that is, the greater the content of opsonins in the blood the greater will be the resistance of the body to bacterial infection. Increase in the number of leucocytes present in the blood is known as *leucocytosis*, and is found as an accompaniment of many disorders.

Besides conveying oxygen to the tissues the blood carries carbon dioxide from the tissues to the lungs for excretion. This carbon dioxide is partly dissolved in the blood, partly present in the form of sodium bicarbonate, and partly combined with the blood proteins. The presence of carbon dioxide in the blood decreases the affinity of the corpuscles for oxygen and hence as the blood travels through the tissues and gathers up carbon dioxide it tends to give up the oxygen which it holds in combination. When it reaches the lungs, on the other hand, the carbon dioxide escapes and the corpuscles have now increased power of taking up oxygen.

CHAPTER XXIII

Excretions of the Body

The waste material leaves the body by way of the lungs, the skin, the feces, and the urine. Through the lungs we lose chiefly the carbon dioxide formed in combustion. The perspiration of the skin consists mainly of water with some inorganic salts and a variety of organic substances present in small traces, and is mixed with the oily secretion of the sebaceous glands. The chief function of the perspiration seems to be to assist in the regulation of the body temperature by supplying moisture which evaporates more rapidly as the body becomes heated, and in so doing removes the excess of heat. It is a well known fact that those persons who perspire freely suffer less in hot weather than those whose sweat-glands are less active. The oily product from the sebaceous glands is probably not to be regarded as excreted waste but as a true secretion which affords a protective covering for the skin and hairs.

In the feces the undigested residue of the food is removed, together with any food material which has not been completely digested or assimilated, and considerable quantities of bacteria from the intestines. Chemical and microscopic examination of the feces in order to detect pathological variations from the normal product is often valuable for diagnostic purposes. For such purpose the patient is usually kept on a simple standard diet such as the following[1] for the experimental period.

[1] Used by Dr. Rehfus at Jefferson Hospital, Philadelphia.

BREAKFAST: 100 grams cream-of-wheat or oatmeal
60 grams toast
20 grams butter
250 c.c. milk

LUNCHEON: Chicken broth with rice
100 grams green vegetable
100 grams mashed potatoes
60 grams toast
20 grams butter
250 c.c. milk

4 O'CLOCK: 250 c.c. milk

DINNER: 150 grams of chopped meat
100 grams green vegetable
100 grams mashed potatoes
60 grams toast
20 grams butter
250 c.c. milk
Stewed fruit

The patient is made to swallow a gelatine capsule containing powdered charcoal or some colored substance which will give a zone of color to the fecal mass at the beginning and end of the experimental period, thus serving to separate the feces of that period from those produced before and after. A similar method is resorted to in nutrition and metabolism experiments in order to be able to calculate accurately the output through the feces in a given period of time.

The urine is the medium through which the waste nitrogenous material passes out of the body. Much can be learned about the process of metabolism from a study of these waste substances, and in particular, variations from the normal course can be fairly easily detected. Consequently careful study of the urine is of great assistance to the physician in diagnosis. For this purpose the whole volume of urine passed in twenty-four hours is usually collected and a sample of the mixed output taken

for analysis in order to avoid misleading results due to temporary variation on account of some superficial change in conditions. Little, if anything, can be learned from tests made on a single sample. In certain pathological conditions it is necessary to examine separately and compare the urine passed during the day and that secreted at night. When this is done, that passed between 8 A.M. and 8 P.M. may be taken as the day sample and that from 8 P.M. to 8 A.M. as the night sample.

One of the most interesting of nitrogenous substances excreted in the urine is *uric acid.* It is produced partly by direct decomposition of certain protein substances, called *purines,* in the food, and partly as a result of the normal metabolism of the tissues. That coming from the food is sometimes distinguished as *exogenous* uric acid, while that produced from the tissues is called *endogenous.* The exogenous acid will naturally vary according to the amount and kind of food eaten, while the endogenous is dependent on the conditions of metabolism. In order to study the production of endogenous acid it is necessary to experiment with subjects from whose diet all food which might give rise to uric acid has been carefully eliminated. Under these conditions it is found that the excretion of uric acid is greatly increased in diseases involving decomposition of the tissues, for instance after extensive burns when the dead tissue is being absorbed, or after the crisis in pneumonia when the exudate of the lungs is undergoing reabsorption.

If through over production or faulty excretion uric acid is allowed to accumulate in the body, the various unpleasant symptoms classed as the *uric acid disorders* follow. Sometimes the excess of acid is deposited as crystals in the kidneys or bladder, a condition known as "gravel;" sometimes the deposit occurs in the joints, giving rise to gouty symptoms; sometimes the excess is

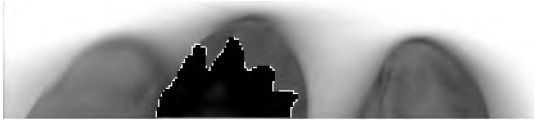
more generally distributed and the condition is known as *lithemia*, and is characterized by various forms of irritation of the nervous system, digestive disturbances, chronic headaches, and occasionally skin affections.

A patient suffering from excess of uric acid should always remain under the supervision of a good physician who will direct his diet among other things, but the general principles to be followed may be indicated. Since the digestive system is usually disturbed by the surplus of acid in the system, all extra strain caused by over indulgence in food should be avoided. All indigestible foods, and in particular any food which tends to be constipating or to increase intestinal putrefaction will be injurious. Alcohol tends to hinder uric acid elimination and is therefore forbidden, except in the case of elderly or debilitated patients for whom stimulation is absolutely necessary, when whiskey is sometimes ordered.

Since uric acid is produced from purines it is of course desirable to eliminate from the diet as far as possible all foods containing these substances. In general meats contain more purines than vegetable food. Sweetbreads, kidney, and liver are particularly rich in purines and should be avoided entirely, while other meats should be used very sparingly if at all. Boiling removes some of the purines and accordingly boiled meats are sometimes allowed in small amounts when roast or fried meats are forbidden. Fish, with the exception of cod, are fairly high in purine content, sardines and anchovies particularly so. Green vegetables are practically purine free, with the exception of spinach, asparagus, peas, and beans, and even these exceptions contain less than meat or fish. Milk, eggs, cheese, nuts, cereals, root vegetables, and fruits may all be regarded as purine free.

Many physicians forbid certain vegetables, not because they contain purines but because they contain an acid,

oxalic acid, which is believed to aggrevate the uric acid disturbances. Among these forbidden vegetables are rhubarb, asparagus, tomatoes, sweet potatoes, and mushrooms. With regard to fruit authorities differ considerably, but nearly all agree in allowing apples, pears, peaches, oranges, and pineapples, and forbidding strawberries on account of the acid which they contain. Where cooked fruits are permitted they are generally stewed without sugar so that they may be less likely to undergo fermentation in the process of digestion. In order to make such fruits palatable they may be sweetened with saccharine, a substance which has a very sweet taste but has no other relation to sugar.



SECTION IV

PRACTICAL MANUAL

Chemicals which Require Special Care in Handling

Potassium and **sodium** react with water with explosive violence. They are best preserved under kerosene oil, care being taken to see that they are well under the surface of the oil. Never throw potassium or sodium into the sink or waste jar. Always handle these metals with forceps, never with the fingers.

Phosphorus oxidizes so rapidly in the air that it may take fire spontaneously. It is therefore always kept under water. This statement and the one in the preceding paragraph should be *carefully noted,* as serious accidents have occurred through careless confusing of the two, and attempting to put sodium in a jar of water. Phosphorus inflicts severe burns if it comes in contact with the skin. The greatest care should be exercised to prevent this but in case of accident remember that the first aid treatment for such a burn is with water in copious quantities, *never oil.*

Bromine is another element which inflicts severe burns. Here again the first treatment in case of accident is with large quantities of water, which should be applied as quickly as possible. Even a slight bromine burn should be shown to a physician at once, as an apparently superficial burn may develop into a severe injury if not properly treated.

All strong **acids** burn, but NITRIC ACID burns are the most serious, being comparable to those of bromine and phosphorus. Even dilute acids eat through cotton rapidly, but they have little effect on woolens except to change the color. This can sometimes be restored by sponging with dilute ammonia.

When concentrated SULPHURIC ACID is mixed with water a large amount of heat is developed. If a little water is added to the strong acid it does not diffuse sufficiently rapidly through the heavy acid and the heat developed may be sufficient to convert the water into steam, with the result that the strong acid is thrown out of the containing vessel. In making a solution of sulphuric acid therefore the acid should always be poured into the water with constant stirring. Under these conditions there is no danger of accident.

Alkalies have a very unpleasant effect on the skin, but do not actually burn except when hot, or in the mouth or eyes. In such a case wash well with water and then with a very dilute solution of acetic acid or vinegar. Alkalies do not destroy cotton, but readily attack wool. They may be carefully neutralized in the fabric with dilute hydrochloric or acetic acid, care being taken to wash out the acid afterwards.

Ether is not only a highly inflamable liquid but very volatile as well, giving off a vapor which when mixed with air is explosive. Ether should therefore never be handled anywhere in the vicinity of a flame, and every care should be taken to prevent the escape of the vapor into a room. So small a thing as the spark caused by turning on an electric switch has been known to cause an explosion in a room filled with ether fumes.

Ordinary GASOLINE and the less common PETROLEUM ETHER should be treated with the same care as ether.

GENERAL DIRECTIONS

Concentrate first on accuracy in following directions, then on efficiency, then on speed.

Cultivate neatness. Pour carefully. Don't leave bottles with drops trickling down their sides. Don't put wet glass rods, stoppers of bottles, and so on down on the table. Don't leave bits of paper, burned matches, etc., lying round. Try not to spill water, but if it should get spilled wipe it up immediately. An untidy worker is a poor worker.

When using a reagent bottle take the stopper between the first and second fingers of the right hand, holding it there while you use the bottle; then replace it without having allowed it to come in contact with anything while it was out of its bottle. This is done to avoid any possible contamination of the reagents through misplaced stoppers.

Use common sense to increase efficiency. Eliminate all unnecessary handling and holding of apparatus, walking round to collect the necessary reagents for a given experiment, and so on.

In heating a liquid contained in a glass vessel be careful that the flame never strikes the surface of the glass above the level of the liquid. This causes unequal expansion and the glass will break. Flasks and beakers should be supported on a wire gauze or asbestos mat in order to prevent too strong heating.

In order to avoid burning the fingers while heating a test tube make a holder by doubling a piece of paper into a strip about an inch wide and folding it around the tube at the top, holding it by the ends (see Fig. 1).

If a liquid is heated strongly at one point, instead of boiling quietly the bubbles of vapor sometimes form irregularly and break so violently that the hot liquid

is often spattered out of the containing vessel. This is known as "bumping." In order to avoid this it must be heated as uniformly as possible, either by keeping the flame moving or by stirring the liquid. Where the heating is done in a test tube the stirring is best accomplished by shaking the test tube constantly with a sharp sideways jerk of the wrist. A little practice will show how this may be done without shaking out the contents. Never attempt to heat a test tube without shaking. Where there is so much liquid that

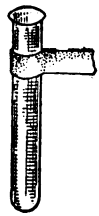


FIG. 1.

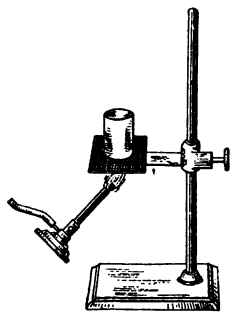


FIG. 2.

it cannot be shaken without spilling, use instead of a test tube a small beaker, the contents of which may be stirred with a glass rod. Bumping may also be avoided by dropping into the liquid two or three small fragments of porcelain (from a broken plate), or better, of pipe-clay. The bubbles form around these and come off more regularly.

Where heating is to be done cautiously it is better to hold the burner in the hand, so that the heat can be regulated by moving it about. When this is done always hold the burner in a slanting position so that the hand will not be directly under the substance which is being heated. In this way there is no danger of being

burned if a flask or beaker should break or its contents should spatter out (see Fig. 2).

For rough estimations of temperature the following standards may be kept in mind. Melting ice is at a temperature of 0°C. Ordinary room temperature averages from 17 to 20°C. Blood heat is about 27°C. 50°C is as high a temperature as the hand will bear comfortably.

To fold filter paper. In order to filter successfully without wasting time it is important to have the filter paper fitted accurately. To accomplish this, fold the round of paper first in two, then in four, and open up the resulting triangular segment so that there are three thicknesses of paper on one side and one on the other.

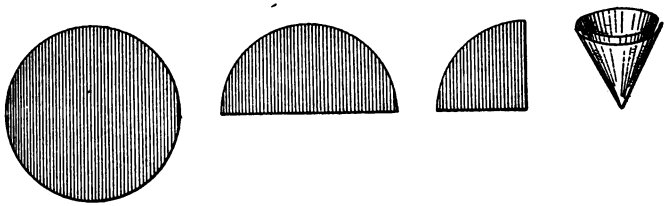


FIG. 3.

If the cone thus obtained does not fit the funnel perfectly it can be enlarged as much as necessary by altering the second fold a little, so that the edges do not come exactly together (see Fig. 3). To test a filter, fill it with water and note the rapidity with which it runs out. It should run in a steady stream which fills the stem of the funnel, with no air bubbles.

To cut glass tubing. With a sharp file make a fairly deep scratch about a quarter of an inch long in the glass. Take the tube in both hands, scratch towards you, fingers outward, thumbs about an inch on either side of the scratch, (see Fig. 4), and bend the glass sharply outward, away from the scratch. If it does not break

readily, cut the scratch a little deeper. Glass so cut will have a sharp edge which should be rounded off. This may be done by rasping it against the flat side of a file, or better, by holding it in the flame until the edge softens and rounds. Remember, never put hot glass down on a cold surface.

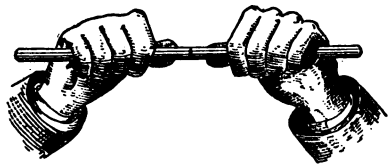


FIG. 4.

To bend glass tubing. Use an ordinary illuminating gas jet or a "fish tail" top on a Bunsen burner, to spread out the flame over a greater amount of surface. Heat the glass gently for a minute, then more strongly, resting it lightly on the fingers of both hands and turning it steadily round and round so that it heats evenly (see

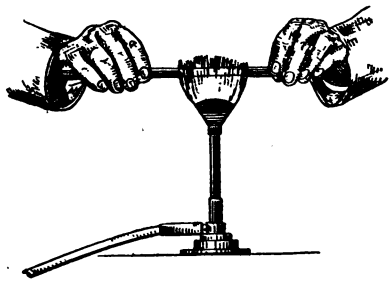


FIG. 5.

(Fig. 5). As it begins to soften the flame takes on a golden glow. Continue heating until it is quite soft, then take it out of the flame, hold it horizontal, and let one end drop by its own weight, merely steadying it with the hand, until it has bent into the angle desired, then hold it steady for a moment until it hardens in that shape.

To bore corks. First soften the cork until it gives a little when pinched between finger and thumb. This can be done by rolling the cork on the floor under the ball of the foot. Select a cork-borer a shade smaller than the tube for which the cork is being bored. Place the cork against a perpendicular surface and cut half way through it with a screwing motion of the cork-borer, being careful to make the hole exactly in the centre, and to keep the borer at right angles to the cork while cutting. Then invert the cork and cut through from the other end until the two cuts meet in the middle of the cork. The core will come out with the borer and the edges of the hole can be smoothed by running the borer through again, first from one end and then from the other. A well-bored cork should have a perfectly smooth cut and fit tightly on the tube, though not so tightly that excessive effort is required to get the tube through. Rubber stoppers may be bored in the same way, but the borer should be kept wet with either alcohol or sodium hydroxide.

In putting rubber tubing or stoppers on glass always wet them a little. This will make the glass slip through much more easily.

In fitting a bent tube into a stopper always hold it between the stopper and the bend, never beyond the bend, otherwise it is liable to break and cut the hand badly. Also, a glass tube should always be held as near as possible to the point where force is applied; that is to say, in the case where a tube is being forced into a stopper, always hold it near the stopper. This reduces the leverage, and hence lessens the danger of breaking. On the other hand, when trying to break glass, as in cutting tubing, the extra leverage aids in the breaking, and therefore the thumbs are placed some little distance apart on either side of the cut.

LABORATORY EXERCISES

1. To Make a Wash Bottle.—A wash-bottle is a flat-bottomed flask (Florence flask) fitted with a two-holed rubber stopper, through which pass two tubes. One of these, the mouth-piece, is between three and four inches long, bent in the middle in an obtuse angle. The other, the delivery tube, is about twelve inches long, and is bent into an acute angle about an inch from one end, this same end being connected by rubber tubing to a short tip drawn out into a fine jet. The longer end goes through the stopper and reaches almost to the bottom of the flask. It is an improvement if, after this tube has been put into the stopper, a slight curve is made in it at the lower end in the same direction as the upper bend. (Fig. 6). If the wash-bottle is partly filled with water and the stopper fitted in place it will be found that by blowing gently through the mouth-piece a fine jet of water is delivered through the tip, and as this tip is flexible, owing to the rubber connection, the stream can be directed to any desired spot. This arrangement is very convenient for washing precipitates in a filter, and if a wash bottle full of water is always on the desk it saves many steps to the sink.

FIG. 6.

To construct. Draw roughly on a piece of paper the angles into which you wish to bend your glass. Take a glass tube and soften it in the flame about two inches from one end. When soft hold it a little above the guide on the paper and bend it gently into the same angle. Let it cool, and cut it off a little more than two inches below the bend. Round both ends in the flame. Bend the other tube in the same way. To make the tip,

heat a piece of tubing in the middle until it is quite soft, then take it out of the flame and draw the two ends gently apart, not too quickly. After it has cooled, cut it where it has been drawn out. Cut off the end to the proper length and round it in the flame. When cool attach it to the long bent tube by an inch and a half of rubber tubing.

2. To Estimate Capacity.—The ordinary test tube holds 25 c.c. Divide it as accurately as you can into five equal parts, making an ink mark to indicate the divisions. Test the accuracy of your eye by measuring 5, 10, 15, and 20 c.c. into it from the laboratory graduate. Correct your divisions if necessary, let the ink dry, and brush over each mark lightly with melted paraffin. Keep this tube for measuring. In the same way graduate a beaker to hold 50, 100, and 200 c.c. In the course of your work train your eye to estimate quantities accurately.

3. Formation of Compounds from Elements.—Make a short coil of copper wire by winding about two inches of wire around a pencil. Put it in a test tube with a pinch of powdered sulphur and heat the two together, gently at first and then more strongly. When the mixture is red hot take the test tube from the flame and notice that the glow increases for a moment or two although the test tube is being cooled. This glow is due to the chemical energy which is set free in the form of heat when the two elements, iron and sulphur, combine. (See p. 16.) After the test tube is cold break it and examine the contents, comparing with the original mixture. How many indications can you find to show that a new substance has been produced? Is there any unchanged iron or sulphur left? Why? How could this be prevented?

Requires: Copper wire.
Powdered sulphur (Flowers of sulphur).
Test tube.

4. Decomposition of a Compound into its Elements. Take about as much mercury rust (red oxide of mercury) as would go on a five-cent piece, and put it in a hard glass test tube. Clamp this tube horizontally in a stand (See Fig. 7) and heat, gently at first and then strongly. At intervals thrust a glowing (not burning) splinter into the tube. What evidences are there of change taking place? Is the gas ("air") in the tube the same at the beginning and at the end of the experiment? Where might the new gas have come from? By what properties would you describe it? The name of this gas is oxygen. After heating for five minutes allow the tube to cool, empty the contents on to a piece of paper and examine them. The silvery liquid is mercury. Where has it come from? Of what elements is mercury rust made up?

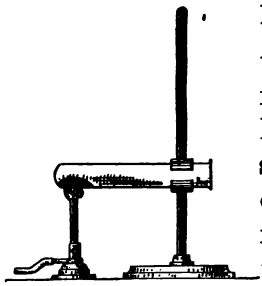


FIG. 7.

The residues should be put into a stock bottle and saved, as mercury and all its compounds are valuable.

Requires: Mercuric oxide.
Hard glass test tube.
Stand and clamp.
Splinters of soft wood.

5. Separation of a Mixture into its Components by Filtration.—Mix thoroughly about a teaspoonful each of salt and sand in a beaker. Heat a little of it in a test tube as you did the iron and sulphur. Is there any evidence of change taking place in this case? To the rest of the mixture add 50 c.c. of water and stir thoroughly. What happens to the salt? Would this action be

hastened or increased by heating the water? Why? (See p. 27.) What is the general method for getting a substance into solution as fast as possible? When the salt is all dissolved, fit a filter paper into a funnel which is supported in a stand. (See Fig. 8). Pour the clear solution carefully through the filter, using a glass rod to guide the stream of water and being careful not to let it flow over the edge of the paper. Collect the portion that comes through (the filtrate) in an evaporating dish, and evaporate it to less than half its bulk. This is most

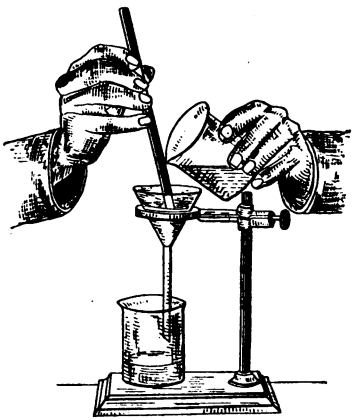


FIG. 8.

quickly done by supporting the dish on a wire gauze over a burner and heating it, first to boiling and then, as it becomes more concentrated, just short of boiling. As the solution becomes concentrated it tends to spatter. The evaporation may be stopped at this point and the solution allowed to cool. It is sometimes more convenient to leave a solution to evaporate slowly by itself at room temperature. In such a case it should be protected from dust by covering it with a paper cone or an inverted beaker which is tipped a little to allow of free circulation

of air. On standing, white crystals separate from the liquid. Dry a few of these with filter paper, grind them to powder and compare with the original salt. How many properties have they in common?

While the salt solution is evaporating add another 50 c.c. of water to the sand, stir well, allow to settle, and pour away the clear liquid. This is known as "washing by decantation." Repeat this once more, and then transfer the sand to the filter. This can be most readily accomplished by washing it out of the beaker and into the funnel with a stream of water from the wash-bottle. Wash the sand in the funnel by pouring water over it until the water coming through has no longer a salty taste.

In this experiment the mixture of salt and sand has been separated into its component parts by taking advantage of the fact that one component is soluble while the other is entirely insoluble. Mixtures of two soluble or two insoluble components can rarely be separated by mechanical means.

Requires: Salt.
Sand.
Beaker.
Funnel and stand.
Filter paper.
Glass rod.
Evaporating dish.
Wash bottle.

6. Distillation. (*Demonstration*).—Arrange apparatus as shown in sketch (Fig. 9). In the flask put water colored with ink or a dye. On boiling the water distils over colorless. If a centigrade thermometer is inserted in the cork of the distilling flask so that the bulb of the thermometer is just below the outlet tube of the flask, the temperature of the boiling vapor may be noted. If this

cork is replaced by another carrying a Fahrenheit thermometer the boiling point of water on the two scales may be compared, and a similar comparison of the freezing points may be made by placing the two thermometers in a beaker of ice and water for a short time.

Requires: Condenser.
500 c.c. distilling flask.
Beaker.
Thermometers (Centigrade and Fahrenheit scales.)
Clamp and stand.
Tripod or stand for flask.

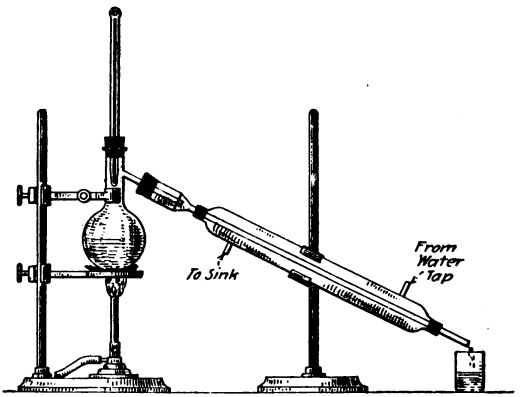


FIG. 9.

7. Decolorization of a Liquid by Animal Charcoal. (*Demonstration*).—In a 500 c.c. flask put 300 c.c. of water colored with a little red ink. Add a handful of powdered animal charcoal, stopper the flask, and shake well for a few minutes. On filtering, the liquid will be found to be colorless.

Requires: Powdered animal charcoal.
500 c.c. flask.
Filter paper.
Beaker.
Funnel and stand.

8. Oxidation. Gain in Weight when Substances are Oxidized. (*Demonstration*).—Heat a porcelain crucible and its lid with a low flame to dry them thoroughly, cool, and weigh to the nearest hundredth of a gram. Polish a piece of magnesium ribbon and fold it back and forth on itself without breaking so that it occupies as small compass as possible, but not so tightly that the air cannot circulate freely through the folds. Nearly fill the crucible with the folded magnesium, taking care that it does not project above the edge of the crucible at any point. Cover the crucible with the lid, leaving just a crack open, and heat it on a pipe-clay triangle. At frequent intervals lift the lid cautiously with a pair of clean forceps. If smoke escapes replace the lid at once. Should there be no sign of reaction increase the heat. Finally, when all the magnesium has lost its metallic lustre and been transformed into the grayish oxide, take the lid off altogether and heat very hot for about five minutes. Cool and weigh.

Requires: Magnesium ribbon.
Porcelain crucible.
Balance and weights.
Pipe-clay triangle and stand.

9. Production of Oxygen by Green Plants in Sunlight. (*Demonstration*). Pass carbon dioxide through a liter or more of water for twenty minutes generating carbon dioxide for the purpose as follows.

In a 500 c.c. flask put a few marble chips or fragments of chalk. Fit the flask with a two-holed stopper, through one hole of which a thistle tube passes down to the bottom of the flask, while the other is fitted with a glass tube bent at right angles. Connect this with rubber tubing to another right angled tube passing through a stopper and down to the bottom of a second flask containing

100 c.c. of water. Through a second hole in the stopper of this flask passes a short right-angled tube which reaches just below the stopper, and which is connected in its turn with a straight tube passing into a large jar (battery jar) of water (see Fig. 10). On pouring dilute

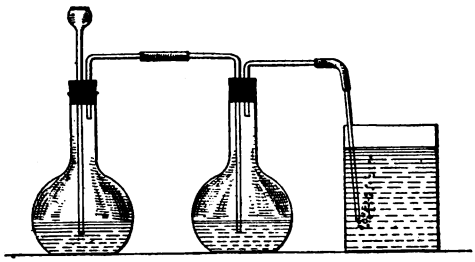


FIG. 10.

(1.5) hydrochloric acid through the thistle tube a stream of carbon dioxide is generated. When the action ceases add more acid, or more marble if necessary.

When the water is saturated, take as many sprays of elodea or myriophyllum (the plants commonly used for gold fish bowls) as will fit into a large (200 × 25 mm.) test tube without crushing. Tie them loosely together and put them in the carbonated water, clipping their stems after they are under water. Put the test tube in the jar, let it fill with water, and invert it with its mouth under the surface of the water, so that the tube remains full of water. Clamp it upright, with the mouth just under the surface, slip the elodea up into the test tube, and place the whole in bright sunlight (see Fig. 11). The plant absorbs carbon dioxide from the water and uses it as food, giving off oxygen as a product of metabolism.

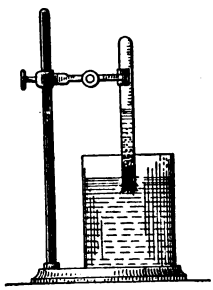


FIG. 11.

When a sufficient amount of gas has collected remove the elodea and test the gas with a glowing splinter. The experiment should not be left running for more than two hours, as by the end of that time the amount of carbon dioxide given off in respiration will tend to exceed the amount of oxygen produced by photosynthesis, and a test for oxygen cannot be obtained. If it is impossible to finish the experiment and make the test during one class period the gas can be kept standing over water until next period, provided the elodea is removed.

Requires: Two flasks with two-holed stoppers, right-angled glass tubes, and rubber connections.
Thistle tube.
Battery jar.
Large test tube (200 × 25 mm.).
Stand and clamp.
Elodea or myriophyllum.

10. **Solution.**—Measure 50 c.c. of water in a beaker, add a small lump of sodium carbonate and note the rate at which it dissolves. Grind an equal amount of sodium carbonate to a fine powder and add it to the same solution. Compare the rate of solution of the powdered substance. Continue to add powdered sodium carbonate, a spoonful at a time, with constant stirring, until no more will dissolve. Then heat the water to boiling and continue to add sodium carbonate until no more will dissolve at that temperature. What effect has temperature on the solubility of sodium carbonate in water? Allow the solution to cool. Explain what happens.

To prepare lime-water. Lime water is a saturated solution of calcium hydroxide, but is best prepared by starting with quicklime, calcium oxide.

To 3 gms. of quicklime add gradually 100 c.c. of water. Let it stand, with occasional stirring, for half an hour.

Allow the suspended parts to settle, decant off the clear liquid and discard it. Add to the residue, which is calcium hydroxide (see p. 39), 1000 c.c. of water, stir thoroughly, leave standing for twenty-four hours, stir thoroughly once more, let the coarser particles settle, and pour the suspension of fine powder ("milk of lime") into a well corked bottle. The solid will settle slowly, leaving the clear saturated solution on top.

Requires: Sodium carbonate.
Quicklime.
Beakers.
"Gem jars," quart size, for preparing limewater.

11. **Osmosis.** (*Demonstration*).—Fill a test tube with a dilute solution of copper sulphate and introduce into this a drop or two of a concentrated solution of potassium ferrocyanide by letting it run in from a dropping tube (or medicine dropper) of which the mouth is just under the surface of the copper sulphate. As the drop comes in contact with the copper sulphate insoluble copper ferrocyanide is precipitated at the boundary of the two solutions, forming a semi-permeable membrane which surrounds the drop. Since the solution outside the little cell thus formed is more dilute than that inside water will tend to pass in and the drop will be seen to swell gradually, until finally, in the course of half an hour or so, the pressure becomes so great that the skin bursts. By reversing this experiment and adding a drop of dilute copper sulphate to a test tube of the more concentrated potassium ferrocyanide we get an illustration of plasmolysis. The water passes out from the drop of dilute solution causing a shrinking and wrinkling of the surrounding membrane.

Requires: Test tube.
Glass dropping tube.
Solution of copper sulphate (7 gms. to 100 c.c.).
Solution of potassium ferrocyanide (20 gms. in 100 c.c.).

12. Study of Carbon Dioxide.—Preparation from carbonates. To a pinch of sodium carbonate in a test tube add a little dilute hyrochloric acid. Repeat with sodium acid carbonate, calcium carbonate (marble), and baking powder.

Set up an apparatus like sketch (Fig. 12), consisting of a test tube fitted with a cork through which passes a tube about four inches long with a right-angled bend in the middle. Connect this by about four inches of rubber tubing to a straight tube passing down to the bottom of another test tube (without a cork). In this second tube put five c.c. of clear lime water, and in the first tube put about half a teaspoonful of sodium carbonate. Add a little dilute hydrochloric acid and quickly replace the cork. Note what happens to the lime water. This is a test for the gas carbon dioxide. (See p. 21.)

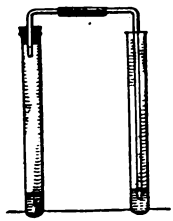


FIG. 12.

Breathe through clear lime water by blowing through a glass tube. Explain what happens.

Pour a little clear lime water into a beaker and leave open to the air for half an hour or more. Explain what happens.

Use the same apparatus as above, but without the lime water. Instead of sodium carbonate use fragments of marble or chalk, and add dilute acid as before. Hold a burning match at the end of the outlet tube. Direct the outlet tube towards the wick of a lighted candle. Explain what happens.

Solubility of carbon dioxide. Allow carbon dioxide to pass into water in a test tube, using the same apparatus as above. After a few moments test the water with clear lime water. Explain the result.

Requires: Sodium carbonate.
Sodium acid carbonate.
Baking powder.
Marble, or chalk.
Dilute hydrochloric acid (1.5).
Lime water.
Test tubes.
Glass tubes and rubber connections.

13. Study of acids.—In each of four test tubes put 5 c.c. of water, and add to one a couple of drops of hydrochloric acid, and to the others an equal amount of sulphuric, nitric, and acetic acids respectively. Dip a glass rod in the solutions and taste each one. Put a drop of each on a piece of blue litmus. To each solution add a few iron filings or tacks. (Any other metal, except copper, would have the same action). Test the gas given off with a flaming splinter. Note carefully what happens both to the gas and to the splinter. This is a characteristic test for hydrogen.

Requires: Acids (Hydrochloric, sulphuric, nitric, and acetic).
Litmus (blue).
Iron filings or tacks.
Test tubes.
Splinter.

14. Study of Bases.—Take small fragments of sodium hydroxide, potassium hydroxide, and sodium carbonate, and dissolve each one separately in about 5 c.c. of water in a test tube. Note the taste and feeling of the solutions and their action on red litmus paper. Do these three substances all belong to the same class? What is the distinction between them? (See p. 39).

Requires: Sodium and potassium hydroxides.
Sodium carbonate.
Litmus (red).
Test tubes.

15. Neutralization.—To 5 c.c. 10 per cent. sodium hydroxide solution in a beaker add about 3 c.c. of dilute (1.5) hydrochloric acid, then continue to add acid, drop by drop, stirring after each addition and testing with litmus paper. In making these tests put the litmus paper on a clean glass surface (the bottom of an inverted beaker will do), take a small drop on a glass rod and touch it to a fresh spot on the paper. When the solution no longer affects red litmus, test it with blue. At neutralization point neither red nor blue litmus is affected. Why?

Requires: Sodium hydroxide (10 per cent.).
Hydrochloric acid (1.5).
Red and blue litmus paper.
Beaker.
Dropping tube and stirring rod.

16. Catalysts and Enzymes. (*Demonstration*).—(*a*) *Manganese Dioxide as a Catalyst.*—Heat 5 c.c. of hydrogen peroxide in a test tube and test the gas evolved with a glowing splinter.

To another 5 c.c. of hydrogen peroxide at room temperature add a pinch of manganese dioxide and test the gas as before.

This is an example of a catalyst bringing about a reaction at room temperature which would otherwise require heat.

Requires: Hydrogen peroxide.
Manganese dioxide.
Test tube and splinter.

(*b*) *Action of the Enzyme of Liver on Fats.*—An extract of liver is made by putting a piece of liver about three inches square through the meat chopper and then grinding it in a mortar with a handful of sand and a little water until it is reduced to a fine pulp, the finer the better.

Add 150 c.c. of water, stir thoroughly for a few minutes, and filter through cheese-cloth.

Melt a bit of butter the size of a small bean, and put it in a test tube with 5 c.c. of the liver extract, shaking the two together until the fat is emulsified. Put a cotton plug in the tube and leave it in the incubator (or in a warm place) at a temperature of not over 30°C. for thirty-six hours. Note the strong unpleasant odor of butyric acid, due to the hydrolysis of the fat. For comparison carry out an exactly similar experiment, using liver extract which has been boiled for a minute. To what is the difference due?

(c) *Precipitation of an enzyme.*—To 25 c.c. of liver extract add double its volume of 95 per cent. alcohol. The enzymes are precipitated (see p. 51), along with much protein material from the liver. Filter the mass, discarding the filtrate, and pour 25 c.c. of water through the filter, collecting it in a small beaker. To this add triple its volume of alcohol and a few drops of strong salt solution. On standing the enzyme is reprecipitated, this time without so much protein contamination.

Requires: Liver extract.
Butter.
95 per cent. alcohol.
Strong sodium chloride solution.
Test tubes and small beakers.
Filter paper, funnel, and stand.

17. Experiments on Fats.—Test the solubility of the fat in water. In alcohol. In ether. What are your conclusions?

Shake up a drop of olive or cod-liver oil with 3 c.c. of water. Does it form an emulsion? Repeat, using soapy water. Result? Repeat using sodium carbonate. Result? (See p. 29.)

Test for glycerine in fat. Mix a drop of glycerine

with a little powdered potassium acid sulphate in a dry test tube and heat. Note the odor cautiously. Repeat, using a fat instead of pure glycerine.

Requires: Fat or oil.
Sodium carbonate solution.
Glycerine.
Potassium acid sulphate.
Test tubes.

18. Preparation of Soap.—Melt about a tablespoonful of fat in an evaporating dish, stirring constantly while heating. When melted, add gradually about 25 c.c. of 10 per cent. sodium hydroxide, beating it into the fat to prevent burning. What reaction takes place? (See p. 91). Then add 50 c.c. of a strong solution of common salt, mix, and leave to cool. When cold lift out the cake of soap, wash well with cold water, melt it with 20 c.c. of water, mixing well, and cool again. What is soap?

Dissolve a little soap in pure (distilled) water. In hard water. (Water not hard may be made so by adding a few drops of lime water to it.) Explain the difference.

To a little soap solution add a few drops of hydrochloric acid. What is the precipitate?

Requires: Fat.
Sodium hydroxide (10 per cent.).
Strong sodium chloride solution.
Evaporating dish.
Hydrocholoric acid.

10. Experiments on Carbohydrates.—(*a*) *Tests for Constituents.* Heat a little sugar in a test tube and note the water given off as the sugar darkens. On strong heating all the hydrogen and oxygen are given off in the form of water and a black mass of carbon is left

Requires: Cane sugar.
Test tube.

(*b*) *Fehling Solution Test for Glucose.* (See p. 98.) Mix about 3 c.c. each of the blue and the colorless parts of the Fehling solution and add a few drops of a solution of glucose. Heat the tube gently and note the change of color. To what is it due? Repeat with cane sugar solution.

To 3 c.c. of cane sugar solution add 3 drops of hydrochloric acid (1.5) and boil for two minutes. Neutralize the acid with a drop or two of sodium hydroxide, and test the solution with Fehling solution as before. Explain.

Requires: Glucose.
Fehling solution.[1]
Hydrochloric acid (1.5).
Sodium hydroxide (10 per cent.).
Test tubes.

(*c*) *Tests with Starch.* Test the solubility of starch in cold water. In boiling water. Make a little "starch paste" by mixing a teaspoonful of starch with 200 c.c. of cold water and boiling for three minutes. Test a little of this starch paste with iodine solution. With Fehling solution. To 1 c.c. add two drops of hydrochloric acid and boil for three minutes. Test again with Fehling solution.

To 1 c.c. of the unaltered starch paste add 3 c.c. of saliva. Put the tube in a beaker of water at a temperature of not more than 30°C. for five minutes. Test with Fehling solution.

Crumble a bit of plain unsweetened cracker (soda cracker) in water and test with Fehling solution. What

[1] Fehling solution is prepared as follows. (Benedict's modification). Dissolve 34.65 gms. of copper sulphate in water and make up to 500 cc. Dissolve 100 gms. of anhydrous sodium carbonate and 175 gms. of Rochelle salt in water and make up to 500 cc. Keep these two solutions separately in rubber-stoppered bottles and mix in equal volumes when needed.

is the chief food constituent of bread, crackers, etc.? Chew a bit of cracker well for at least a minute and test with Fehling solution. What has happened? To what is it due?

Action of diastase with starch. To the remainder of the starch solution add a diastase capsule and keep in warm water at as nearly 50°C. as possible. It must not be allowed to warm up above 65°. At intervals of five minutes test a few drops with iodine solution. When the iodine no longer shows starch present, test with Fehling solution. What has taken place, and to what is it due?

Requires: Diastase capsules.
Starch.
Fehling solution.
Iodine solution. (Iodine crystals dissolved in a dilute solution of potassium iodide.)

(*d*) *Fermentation.*—Mix 5 c.c. of concentrated glucose solution ("corn syrup" is very satisfactory) in a test tube with quarter of a yeast cake made into a paste with 5 c.c. of water. Close the test tube by means of a cork with a right-angled delivery tube, connected by rubber with another tube dipping below the surface of a little clear lime water in a second test tube which is protected from the air by a cotton plug. (See Fig. 12.) Leave standing until next exercise.

At the next period note the appearance of the lime water. What does this indicate? Note the alcoholic odor of the fermenting liquid. Filter the liquid, add two or three small crystals of iodine, and warm gently for a few moments. Cool, and add sodium hydroxide until the color of the iodine disappears. On standing, yellow crystals of iodoform are deposited and may be recognized by their characteristic odor. The formation

of iodoform is a test for the presence of ethyl alcohol. (See p. 80.)

Requires: Glucose.
Yeast.
Lime water.
Iodine crystals.
Sodium hydroxide solution.
Test tubes and connecting tubes.

(*e*) *Test for Sugar in Milk.*—To 10 c.c. of milk in a beaker add an equal volume of water, and warm gently to between 50° and 60°C. Now add not more than 1 c.c. of 5 per cent. acetic acid. By this treatment the casein of the milk is precipitated. Stir thoroughly, allow to settle, decant the clear liquid through a filter paper, and test the filtrate with Fehling solution.

Requires: Milk.
Acetic acid (5 per cent.).
Fehling solution.
Beaker, funnel, and test tube.

20. Tests on Proteins.—Put a drop of strong nitric acid on a little protein material. Note the color produced.

Heat a little protein with 2 or 3 c.c. of Millon's reagent. Note the color produced.

Heat a little protein with "biuret reagent" in an evaporating dish. Note the color change.

Test for the Constituents of Protein. (*a*) *Nitrogen.*—To a small quantity of dry gelatine add over ten times its amount of soda-lime, and grind the two together in a mortar till well pulverized. Transfer the mixture to a dry test tube and heat. Test the gas given off with red litmus. Note the pungent odor of ammonia (a little disguised by the burning organic matter).

(*b*) *Sulphur.*—Mix about a half teaspoonful of dry powdered egg albumin with twice its bulk of "fusion

mixture" in a porcelain crucible. Set the crucible on a pipe-clay triangle (see Fig. 13) and heat cautiously. As there is a disagreeable fume produced in this process the experiment should be carried on in a draft cupboard or near an open window through which the fumes can escape. As soon as the action begins in the crucible, take away the flame until it has moderated. Continue the fusion until the product is practically colorless, then allow the fused mass to cool. Fill the crucible with water and if necessary warm gently to dissolve the fused material. Pour the solution into a small beaker and acidify with a few drops of hydrochloric acid. Transfer to a test tube, filtering if not perfectly clear, and add a little barium chloride. A white precipitate of barium sulphate indicates the presence of sulphur.

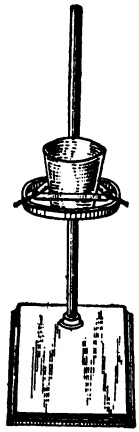


FIG. 13.

Coagulation. To 20 c.c. of egg white add four times its bulk of water. Heat, and note temperature of coagulation.

Requires: Protein material. (Either coagulated egg white of the casein of milk can be used for the color tests.)
Gelatine. Powdered egg albumin.
Powdered egg albumin.
Nitric acid.
Millon's reagent.[1]
Biuret reagent.[2]

[1] Millon's reagent is prepared as follows. Dissolve 200 gms. of mercury in its own weight of concentrated nitric acid (in draft cupboard, on account of fumes given off). Treat the solution with twice its volume of water, allow to stand, and decant the clear solution from the sediment which forms.

[2] Gies' biuret reagent is prepared as follows. To 500 c.c. of 10 per cent. sodium hydroxide solution add 12.5 c.c. of 3 per cent. copper sulphate solution.

Soda lime.
Fusion mixture (1 part Na_2CO_3 and 3 parts KNO_3 powdered and mixed together).
Hydrochloric acid.
Barium chloride solution (112 gms. per litre).

21. Hydrolysis of Protein (*Demonstration*).—Put the whites of three eggs in a flask and heat in boiling water until coagulated. Add 300 c.c. of water in which has been dissolved 15 c.c. of concentrated sulphuric acid, and leave on a water bath (or on a radiator) until the protein material has been decomposed, leaving only a fluffy white insoluble residue. Filter this off and add to the filtrate 300 c.c. of 20 per cent. barium hydroxide. Then add more carefully about 25 c.c. more, testing with litmus after each addition until neutral. Should the neutral point be passed add a few drops of 5 per cent. acetic acid until the solution gives a very faint acid reaction. Allow it to stand over night, or until the precipitated barium sulphate has settled down to the bottom of the beaker. Decant the clear solution through a filter, finally transferring the precipitate to the funnel. Collect the filtrate in an evaporating dish and evaporate it to half volume on a wire gauze, and then on a water bath to about 10 c.c. To the syrupy residue after evaporation add 95 per cent. alcohol until no more precipitate forms, stirring continually. The sticky precipitate consists of peptones and proteoses (see p. 112). Gather this precipitate together as much as possible on the side of the dish, warm the solution gently, and decant through a dry filter paper. The filtrate contains amino acids (tyrosine and leucine) which will separate in crystalline form if the solution is allowed to evaporate slowly.

Requires: Egg white.
Barium hydroxide (20 per cent. solution).

95 per cent. alcohol.
500 c.c. flask.
750 c.c. beaker.
Evaporating dish.
Funnels and stand.

22. Experiments on Blood. (Demonstration.)—(*a*) *Clotting of Blood.*—Leave 300 c.c. of freshly drawn blood in an open beaker until a solid clot has formed. Then cover and allow to stand for at least twenty-four hours. The clot shrinks and the yellowish serum separates from it.

(*b*) *Defibrination of Blood.*—Stir 300 c.c. or more of freshly drawn blood vigorously for five to eight minutes. Collect the solid which separates and wash it well with water. This is fibrin (see p. 131). Test it for protein. The residue from which the fibrin has been removed is called defibrinated blood.

(*c*) *Reactions of Haemoglobin.*—Take 10 c.c. of defibrinated blood in each of three test tubes. Shake one well with air so that it can take up as much oxygen as possible. The haemoglobin is converted into oxyhaemoglobin. To the second add a few drops of Stokes reagent.[1] This is a strong reducing agent and reduces the oxyhaemoglobin present to haemoglobin. Pass illuminating gas through the third tube. Illuminating gas contains carbon monoxide, which unites with the haemoglobin (see p. 133). Note the difference in color in the three tubes.

(*d*) *Laking of Blood.*—To a few c.c. of defibrinated blood add distilled water little by little and note the effect. If a microscope is available examine slides of the blood before and after laking (see p. 134).

[1] Stokes reagent is prepared as follows: Dissolve 3 gms. of ferrous sulphate in 25 c.c. of cold water. Dissolve 2 gms. of tartaric acid in the same amount of water. Mix the two solutions and add 50 c.c. water. Just before using add strong ammonium hydroxide until the precipitate first formed just dissolves.

CHEMICALS AND APPARATUS REQUIRED FOR A CLASS OF TEN STUDENTS

(Items marked # are used for demonstration or for making up a quantity of solution, and need not be increased for a larger number of students.)

CHEMICALS

Amount	Name
1 lb.	Acid, Acetic#
	Hydrochloric#
	Nitric#
	Sulphuric#
1 liter	Alcohol, 95 per cent.#
2 oz.	Animal charcoal (powdered)#
¼ lb.	Barium chloride#
¼ lb.	Barium hydroxide#
2 oz.	Calcium oxide
½ lb.	Calcium carbonate (marble chips)
¼ lb.	Copper sulphate (cryst.)#
3 ft.	Copper wire (medium size)
15	Diastase capsules
¼ oz.	Egg albumin powdered
3 gms.	Ferrous sulphate#
1 oz.	Gelatine (powdered, preferably)
1 oz.	Glucose (corn syrup)
¼ oz.	Glycerine
20 c.c.	Hydrogen peroxide#
¼ lb.	Iron filings or tacks
1 oz.	Iodine
1 roll	Magnesium ribbon
1 gm.	Manganese dioxide, powdered#
¼ lb.	Mercuric oxide (Red oxide of mercury)
½ lb.	Mercury#
1 oz.	Potassium acid sulphate
20 gms.	Potassium ferrocyanide#
¼ lb.	Potassium hydroxide#
5 gms.	Potassium iodide
¼ lb.	Potassium nitrate
½ lb.	Rochelle salt (sodium potassium tartrate)#
1 lb.	Soda lime

Amount		Name
100	gms.	Sodium carbonate (Anhydrous)#
2	lbs.	Sodium carbonate, crystalline (Washing soda)
1	oz.	Sodium bicarbonate
1	lb.	Sodium chloride
1	lb.	Sodium hydroxide
½	lb.	Starch
½	lb.	Sugar (cane)
1	oz.	Sulphur (Flowers)
2	gms.	Tartaric acid#

APPARATUS

Balance and Weights.—For student use the small horn pan balances (Fig. 14) will be found very satisfactory. They are sentitive to 1/100 of a gram. For weighing larger quantities, as in making up solutions for laboratory use, the Harvard trip balance (Fig. 15) is more convenient. Metric weights from centigrams up to 200 gms. will be required.

Amount	Name
1	Battery jar# (for exp. 9). Plain glass jar (Fig. 16). Height 8 in., diameter 6 in.
30	Beakers (Fig. 17). 350 c.c. capacity
5	500 c.c. capacity#
2	750 c.c. capacity#
2 doz.	Bottles # ½ pt. (500 c.c.). Narrow mouth.
½	6 oz. Salt mouth.
½	4
1 doz.	Bunsen burners (Fig. 18). Where acetylene gas is used a special form of burner is required. If the laboratory is not supplied with gas, small alcohol lamps may be used satisfactorily.
10	Clamps and holders, small size (Fig. 20), for exp. 4.
1	large (Fig. 21), for exp. 6.#
1	Condenser, Liebig's, (Fig. 22) for exp. 6.#
1 set	Cork borers.
2 gross	Corks, No. 6.
15	Crucibles and lids (Fig. 23). Capacity about 12 c.c.
12	Evaporating dishes (Fig. 24). About 200 c.c. capacity
3	1½ liters, for making solutions, etc.#
1 pkg.	Filter paper, diameter, 5 in.
1	8 in.#
18	Flasks, flat bottomed or Florence (Fig. 25) 1000 c.c. capacity.
3	500 c.c. for exp.9.#
2	distilling (Fig. 26) 500 c.c. for exp.6.#
10	Forceps
15	Funnels (Fig. 27) diameter 3 in.
3	5 in.#
10	Fish-tail or wing tops for burners (Fig. 19)
10	"Gem jars," quart size, for preparation of lime water.
2 doz.	Glass rods, length 5 in.
1 lb.	Glass tubing, outside diameter about 7 mm.

Amount		Name
3		Graduates 100 c.c.#
2		1000 c.c.#
12		Wire gauzes, 5 × 5 in. (Nichrome wire, although more expensive in the beginning wears better than the cheaper varieties.)
2	pkgs.	Litmus paper, red.
2		blue.
10		Mortars and pestles (Fig. 28) diameter 10 cms.
1		16 cms.#
10		Pipe-clay triangles
10		Ring stands (supports), (Fig. 29). Small size, with two rings.
2		Large size, with four rings.#
1	lb.	Rubber stoppers, two-hole, no. 6
6	ft.	Rubber connection tubing, inside diameter, 3/16 in.
6	ft.	condenser 1/4 in.#
24	ft.	burner 5/16 in.
3		Test tubes (Fig. 30). Size 8×1 in. for exp. 9.#
6	doz.	Test tubes. Size 6×5/8 in.
1	doz.	Test tubes, hard glass, 6×5/8 in. for exp. 4.
1		Thermometer, centigrade scale.#
1		Fahrenheit.#
1		Thistle tube (Fig. 31).#
10		Water baths. A water bath is the chemist's equivalent for a double boiler, and is a vessel in which water can be boiled and in which a smaller vessel can be placed. For small amounts of material two beakers serve very well, the larger one containing water, and the smaller the substance to be heated. For larger amounts a kitchen sauce pan can be used as the outer vessel. The dealers supply copper rings which fit the top of the sauce pan making it possible to adjust the size of the opening to the size of the vessel to be heated.
10	sets	Rings for water baths, in sets. Outside diameter of largest ring at least 7 in.

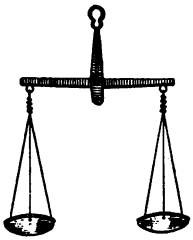


FIG. 14.

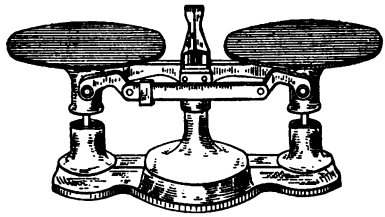


FIG. 15.

FIG. 16.

FIG. 17.

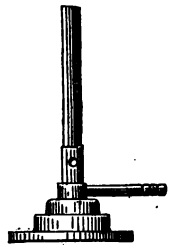


FIG. 18.

FIG. 19.

FIG. 20.

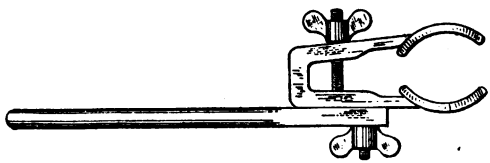


FIG. 21.

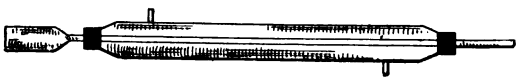


FIG. 22.

FIG. 23.

FIG. 24.

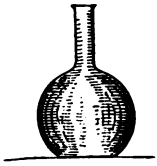


FIG. 25.

FIG. 26.

FIG. 27.

FIG. 28.

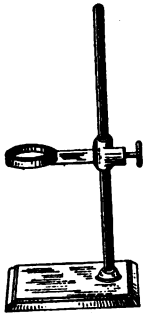


FIG. 29.

FIG. 30.

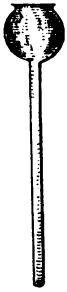


FIG. 31.

List of the Elements, their Symbols, Atomic Weights, and Valence.

INTERNATIONAL ATOMIC WEIGHTS

1920

Name	Symbol	Atomic weight	Principal valence
Aluminum	Al	27.1	3
Antimony, stibium	Sb	120.2	3 or 5
Argon	A	39.9	0
Arsenic	As	74.96	3 or 5
Barium	Ba	137.37	2
Bismuth	Bi	208.0	3 or 5
Boron	B	10.9	3
Bromine	Br	79.92	1
Cadmium	Cd	112.40	2
Cæsium	Cs	132.81	1
Calcium	Ca	40.07	2
Carbon	C	12.005	2 or 4
Cerium	Ce	140.25	4 or 3
Chlorine	Cl	35.46	1
Chromium	Cr	52.0	2, 3 or 6
Cobalt	Co	58.97	2 or 3
Columbium, niobium	Cb	93.1	3 or 5
Copper	Cu	63.57	1 or 2
Dysprosium	Dy	162.5	3
Erbium	Er	167.7	3
Europium	Eu	152.0	3
Fluorine	F	19.0	1
Gadolinium	Gd	157.3	3
Gallium	Ga	70.1	3
Germanium	Ge	72.5	4
Glucinum, beryllium	Gl	9.1	2
Gold, aurum	Au	197.2	1 or 3
Helium	He	4.00	0
Holmium	Ho	163.5	3
Hydrogen	H	1.008	1
Indium	In	114.8	3
Iodine	I	126.92	1
Iridium	Ir	193.1	3 or 4
Iron, ferrum	Fe	55.84	2 or 3
Krypton	Kr	82.92	0
Lanthanum	La	139.0	3
Lead, plumbum	Pb	207.2	2 or 4
Lithium	Li	6.94	1
Lutecium	Lu	175.0	3
Magnesium	Mg	24.32	2
Manganese	Mn	54.93	2, 4, 6 or 7

Name	Symbol	Atomic weight	Principal valence
Mercury, hydrargyrum	Hg	200.6	1 or 2
Molybdenum	Mo	96.0	3, 4 or 6
Neodymium	Nd	144.3	3
Neon	Ne	20.2	0
Nickel	Ni	58.68	2 or 3
Niton, Ra emanation	Nt	222.4	
Nitrogen	N	14.008	3 or 5
Osmium	Os	190.9	2, 3, 4 or 8
Oxygen	O	16.00	2
Palladium	Pd	106.7	2 or 4
Phosphorus	P	31.04	3 or 5
Platinum	Pt	195.2	2 or 4
Potassium, kalium	K	39.10	1
Praseodymium	Pr	140.9	3
Radium	Ra	226.0	2
Rhodium	Rh	102.9	3
Rubidium	Rb	85.45	1
Ruthenium	Ru	101.7	3, 4, 6 or 8
Samarium	Sa	150.4	3
Scandium	Sc	44.1	3
Selenium	Se	79.2	2, 4 or 6
Silicon	Si	28.3	4
Silver, argentum	Ag	107.88	1
Sodium, natrium	Na	23.00	1
Strontium	Sr	87.63	2
Sulphur	S	32.06	2, 4 or 6
Tantalum	Ta	181.5	5
Tellurium	Te	127.5	2, 4 or 6
Terbium	Tb	159.2	3
Thallium	Tl	204.0	1 or 3
Thorium	Th	232.15	4
Thulium	Tm	168.5	3
Tin, stannum	Sn	118.7	2 or 4
Titanium	Ti	48.1	3 or 4
Tungsten, wolframium	W	184.0	6
Uranium	U	238.2	4 or 6
Vanadium	V	51.0	3 or 5
Xenon	Xe	130.2	0
Ytterbium	Yb	173.5	3
Yttrium	Yt	89.33	3
Zinc	Zn	65.37	2
Zirconium	Zr	90.6	4

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